

EXPERIMENTAL EDITION

SCHOOL CHEMISTRY

CLASS 6

TEACHER'S GUIDE

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NATIONAL COUNCIL OF EDUCATIONAL
RESEARCH AND TRAINING

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SCHOOL CHEMISTRY

CLASS 6

TEACHER'S GUIDE

Prepared by

N C E R T Chemistry Study Groups



Under a grant from

**THE NATIONAL COUNCIL OF EDUCATIONAL
RESEARCH AND TRAINING**

NEW DELHI

PREFACE

From time to time new textbooks have been written in chemistry, mainly to present additional facts brought to light by researches. In almost all these books the facts are given as a series of statements. It is usual to write, for instance, that 'acids are substances having a sour taste, are corrosive to the skin, turn blue litmus red, contain hydrogen which can be liberated by metals like zinc, iron, etc'.

The pupils read and memorise these 'important' facts and practically repeat the statements in an examination. It is only at a much later stage, if at all, that they may be called to verify some of the statements in the laboratory. The thrill and joy of discovery is lacking, as they have already been told what to expect. Such a study becomes dull and uninteresting and is hardly different from a study of remote events in history. This is how most of us may have learnt science.

Contrast this approach with the one where the pupils first do things by themselves in the laboratory, record their observations, draw conclusions and then try to find some reasonable explanation for the observed facts. For instance, the pupils may be given a few unknown solutions. Some of them are acids, some bases and the rest neither. The pupils are advised to try the action of these solutions on litmus, iron filings and sodium carbonate, etc. They record their observations and find for themselves, when told what the solutions are, the characteristic properties of acids and bases and how these differ from one another and from others. Learnt this way, knowledge is more easily retained, without the strain of memorization. There is also the thrill of finding something new by themselves.

This is the new approach which lays emphasis on pupils' experimentation and learning through their own observations. The topics covered in the text follow, and not precede, the experiments in the laboratory manual. The teacher's role is mainly to guide in laying down the general procedure and help in organising the laboratory work. No statements or facts are expected to be made which the pupils are supposed to accept and remember. Doubtless, certain topics cannot be experimentally verified at such an early stage by the pupils. Such topics may be explained with the aid of charts, diagrams and models and, wherever available, by means of filmstrips. However, topics of a theoretical nature are discussed only to the extent necessary for an intelligent understanding and correlation of the observed facts.

Teaching chemistry according to this new approach requires considerable effort on the part of the teacher. The members of the Study Groups hope that the teacher's guide will prove to be of some help to the teacher in achieving this objective. Though the teacher is expected to keep to the scheduling of the units, he has considerable freedom in organising the detailed work. Although certain experiments have been suggested, he may design new experiments to fulfil the same objectives. Without experimental basis the course will not be meaningful.

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TO THE TEACHER

In view of the great strides that chemistry has made in recent years, it has become necessary to expose the students to the modern ideas and trends in chemistry (even at the school level) so that they keep pace with modern developments. To make the study of the subject interesting and intelligible, a new approach is attempted in which the pupils are made to learn the basic principles of chemistry through experimentation instead of the traditional way of teaching which leads to memorisation of facts. This approach is believed to promote a sense of experimentation and inquiry in the young minds. Besides, it enables them to develop sufficient skill in the laboratory techniques and appreciate the need for careful observation.

May be it is the first time a project of this type is placed before you for teaching. The success of this depends on your effort. The facts learnt in the laboratory and through demonstration experiments should form the subject matter of the text. Before each laboratory session, careful planning and preparation is necessary for successful conduct of the practical class. Gather all the materials required (apparatus, chemicals and solutions) well before the class commences. The materials required for each experiment are mentioned in the laboratory manual. Supply the pupils only the required amounts of the materials to avoid wastage. A list of apparatus and the amounts of chemicals required for a class of 40 pupils for all the experiments of the class is given in this guide so that the laboratory is equipped with all the required chemicals and apparatus.

It is desirable that, as far as possible, the experiments are done by the pupils individually. Try to use wherever possible improvised apparatus, made of materials locally available. The pupils should be taught the correct way of handling the materials. It is important that they are not told as to what will happen before hand. This would spoil their keen observation, curiosity, and spirit of inquiry. Go round the class to see whether the pupils are doing the experiments properly and check if they missed any observation(s). See that the pupils maintain a separate note book for recording observations. Check the answers of the pupils soon after the completion of each experiment and make them repeat the experiment if the observations are incorrect. (Answers to questions under each experiment are given in this guide for your guidance).

To supplement the experimental work and to cover as many concepts of the syllabus as possible, some demonstrations are described in this guide. The equipment and chemicals needed for them are also listed separately. They should be procured and the demonstrations performed well before they are shown in the class. Under each demonstration, a few questions are suggested. Many such questions may be asked during the conduct of the demonstration to direct the pupils attention to important points to be taken note of.

Follow as far as possible the schedule given in each unit. Needless to say it is very useful for planning the lesson. One lesson consists of two consecutive periods of forty minutes duration each. At the end of each unit a quiz is given. It is only suggestive. You may prepare many such quizzes for examining the pupils.

It is hoped that the instruction material of this guide will be helpful for successful handling of the class. The background information is meant for extending the horizons of your knowledge and should not be taught to the pupils. The teachers' guide is meant for your use and should not be placed in the hands of the pupils. Any helpful suggestions and comments are welcome.

CONTENTS

<i>UNIT</i>	<i>Page</i>
1. CHEMISTRY AND THE MATERIAL WORLD ...	1
2. STRUCTURE OF MATTER	24
3. SYMBOLS, FORMULAE AND EQUATIONS ...	41
4. SOLUTIONS	52

Chemistry and the Material World

INSTRUCTION MATERIAL

Introduction

Man, like other living beings, reacts to his surroundings. Curiosity on his part has been the principal motivating force for the great strides that human civilization has made. Curiosity induces him to observe more closely the objects and the phenomena taking place around him. Observation sets him thinking. Thinking helps develop a rational attitude. These sequences lead to the development of what may be called a "scientific outlook".

Nature produces a wide variety of materials. A study of these materials and attempts to produce them in laboratory is one of the important activities of man. In some cases, he has been successful in preparing them in laboratory and some of them excel in quality those obtained from nature. Behind every such effort, is his incessant search for a happier, more comfortable and meaningful life; besides of course, the thrill and excitement of "discovery".

For the child, observation begins at home. Initially, he takes things as they are — without question. Later, what his parents say about things around him is good enough. But as he grows, he is perhaps not satisfied with these answers. With a far more keen sense of observation than the grown-ups — perhaps because he is "seeing" a number of things for the first time — he begins to wonder. Naturally he is interested in knowing about what he sees. His mind is full of questions — why is this? how is that? — and so on. If he is guided and helped in knowing, preferably by personal experience, the answers to his questions, the thrill and excitement of knowing is all the more great.

It is the intent and purpose of the present course to help the growth of knowledge by encouraging the pupil to "observe" and "wonder why", because scientific activity begins with careful observation. The teacher has a great responsibility here. He should be instrumental in helping the pupil to know what to observe. In the first instance, observation can begin by looking at the myriad objects and phenomena around us in everyday life. The common place objects and happenings — a burning log of wood, the smoke coming out of a factory chimney, the steam from boiling water — can be brought up for consideration, to help the pupil get curious about why and how such things occur. Explanations are sought, analogies drawn, and models constructed. The simpler phenomena observed can be repeated in the laboratory — a place where conditions of reactions are easily controlled. Thus "experimentation" begins. Observations from experiments are recorded, information is gathered and sifted to find why regularities exist. Findings are made known to others. Experimental results help in the growth of science.

In this unit the study of chemistry begins with an attempt to find what constitutes chemistry and its importance in daily life. An observation of common materials around us is made. The various states of matter and the changes that matter undergoes are discussed. Emphasis is laid at a later stage on the important fact that in spite of the apparent differences in the state of aggregation, appearance, colour, shape, size and other properties, the fundamental building blocks of every type of matter are the same.

It is important to realise that the approach to the study of chemistry in this course is intended to be different from the methods followed hitherto. Chemistry is essentially an experimental science. The present course is based on **experimentation, observation and generalisation**. This should always be borne in mind while presenting the course. It is suggested :

- (i) to insist on independent observations by pupils after doing an experiment or watching a demonstration,
- (ii) to encourage the pupils to organise the facts and to draw certain generalised conclusions on the basis of the observations,
- (iii) to develop concepts through experiments and visual aids such as charts, diagrams, and film strips,
- (iv) to draw conclusions from experimental facts and not make general statements.

While dealing with the changes in matter, the traditional method is to start with the definitions of physical and chemical changes. The two types are compared and contrasted in a tabular form supported by suitable examples. The categorical statements may or may not be followed by experiments. In such a method, the pupils take the statement of the teacher as gospel and try to memorise it. They may even successfully reproduce it at the examination but the faculty of creative thinking and logical reasoning in them is not developed. In the

present approach we place before the pupils a few selected materials, allow them to handle them and draw their attention to some of the characteristic properties of the materials. The steps of the experimental procedure and the questions are so designed as to make them observe for themselves the various changes the materials undergo. In the postlab discussion, gather all the observations and help them in organising the information and drawing logical conclusions. The pupils should not be given the definition of either physical or chemical changes but should be led to know about them from observation.

Outline

1. What is chemistry? (1.1)
2. Importance of chemistry in daily life (1.2).
3. Matter and the changes it undergoes (1.3).

Concepts

1. Chemistry deals with substances, their composition and the changes they undergo.
2. Chemistry helps us in different spheres of life.
3. Substances exist in different states under different conditions.
4. Substances may be classified into mixtures, solutions, compounds and elements.
5. Substances undergo changes; in some cases new substances are formed.
6. The total mass of substances remains constant in any of these changes.

SCHEDULE

Lesson	Experiment	Postlab	Demonstration	Text covered
1	1 and 2	Discussion	—	1.1
2	3 and 4	Discussion	—	1.1
3	—	—	Aids & Activities	1.2
4	—	—	I, II and III	1.3
5	—	—	IV A and IV B	1.3
6	5	Discussion	IV C	1.3
7	—	—	V A and V B	1.3
8	—	—	VI	1.3
9	—	—	—	Review and quiz

Development

1.1 What is chemistry?

The content of this section is to be developed on the basis of the laboratory experiments 1, 2, 3 and 4. One or two days before starting experiment 1, ask the pupils to get one or two common materials they come across. No specific mention be made about any particular thing. The following materials are suggested as examples of common substances. The teacher might add a few more things so as to include the materials common in a particular locality. Six substances may be given to each pupil or group of pupils. They must be representative samples of each of the states — solid, liquid and gaseous. Some may be coloured and some may not be. Sand and salt may be given to all.

Solids	Liquids
Copper plate	Petrol
Sand	Water
Pith cork	Vinegar
Common salt	Honey
Aluminium sheet	Kerosene
Wax candle	Coconut oil
Pencil	Ink
Iron nails	Spirit
Sugar	Groundnut oil
Charcoal	
Glass plate	Gases
Rubber	
Match sticks	Air-filled balloon
Wooden block	
Rock	
Sulphur	

On the first day after acquainting the pupils with simple laboratory equipment listed in experiment 1 and the care in handling them, they may be asked to start the experiment

Experiment 1 Study of some common substances

Time required: 60 minutes

In the postlab discussion, ask selected pupils to describe their substances and what properties they have observed. Note these on the black board. Ask which other pupils had any of these substances and whether they also have observed similar properties. With discussion, draw out the correct properties of each substance. Instruct the pupils to correct their observations wherever they

had gone wrong. Give encouragement to pupils who had noticed additional properties and used 'finer adjectives' to describe the properties, for example, a white solid, a shining solid, a soft solid etc.

After this discussion, with the help of pupils, classify on the board all these substances into solids, liquids and gases; soluble and insoluble; light and heavy. Conclude the discussions by bringing out the idea that there are a variety of materials around us. These materials resemble in some respects and differ in other properties. They can exist in different states such as solids, liquids or gases and materials can be sorted out according to their properties.

Having made them familiar with a variety of materials, the next idea to be developed is that materials undergo changes when heated. The convenient source of heat in a school laboratory is a spirit lamp. Experiment 2 familiarises the pupils with the proper use of a spirit lamp.

Experiment 2 Study of a spirit lamp

Time required : 20 minutes

Keep an assembled spirit lamp in front of the pupils throughout the experiment. Instruct the pupils to assemble the spirit lamp following the steps shown by you. Show them how to wet the wick and to arrange it properly for obtaining a good flame. Precautions (i) and (ii) in the laboratory manual are necessary because spirit is inflammable. The pupils must not be allowed to blow off the flame. They should use the cap to put out the flame.

Questions and Answers

2. How do you get a good flame with a spirit lamp?

A good flame is obtained by the proper adjustment of the length and spread of the wick.

3. What materials are used up when the lamp burns?

Spirit.

Experiment 3 Behaviour of substances on heating

Time required : 60 minutes

This experiment has been designed to introduce to the pupils the changes that substances undergo on heating. Since this is the first experiment involving heating, tell them that a dry test tube must be used whenever a substance is to be heated.

Show them

- (i) the correct way of introducing a solid substance into the test tube using a paper boat. The solid should not stick to the side of the test tube and obstruct observation.

- (ii) how to hold the test tube while heating it. (The test tube should be slightly inclined and its mouth should be pointing away from the experimenter). Tell them to use a test tube holder or folded paper to hold the test tube.

After completing any experiment let the pupils clean their test tubes and ask them to form a habit of keeping these clean test tubes inverted in a test tube stand, so that dry test tubes will always be available for any experiment.

Instruct the pupils to be alert and make observations with care. Ask them to take a small quantity of the substance and heat it for not more than 3-4 minutes and observe carefully throughout this period. Changes in state or colour of the substance, evolution of any gas and any deposit on the cooler parts of the test tube are to be recorded. The colour and odour of any gas formed and of the residue left behind should be noted. Various changes that the substances in experiment 3 undergo are mentioned below. **Do not reveal the changes to the students beforehand.**

On heating

- (a) **Nichrome*** wire glows. It regains its original appearance on cooling. No new product is formed.
- (b) **Magnesium ribbon** starts burning. It continues to burn with a dazzling light. The product crumbles to a white powder which is different from the original metal ribbon.
- (c) **Ammonium chloride** — vapourises and condenses as a white sublimate on the cooler parts of the test tube. No new substance is formed.
- (d) **Copper sulphate** (blue crystals) changes into a white powder. A colourless liquid (water) condenses on the cooler parts of the test tube. Show them that on adding a few drops of water to the white powder after cooling, the blue colour reappears.
- (e) **Lead nitrate** melts with a crackling sound. A reddish brown gas (nitrogen dioxide) is evolved. The colour of the gas is best seen by viewing down the test tube. A yellow powder (lead oxide) is left in the tube. The reddish brown gas and the yellow powder — the products observed in this experiment — are different from lead nitrate.

Question and Answer

Do all the substances behave similarly when heated ?

No.

* Pieces of nichrome wire may be obtained from used up heating elements

In the postlab discussion of experiment 3 collect all the observations made by the whole class with each substance. This will give an opportunity to the pupils to know if they have failed to make some observation. However, never discourage or cut short even an obviously incorrect observation being reported. Point out the fallacy in the observation and reason out what might have been responsible for such an observation. If the whole class has missed some important observation ask the pupils to repeat the experiment and put suitable questions while they are observing so as to lead them to the missing observation(s).

Direct the attention of the pupils to the changes they have observed and which are noted on the board. Impress upon them that it is these changes that substances undergo that lead to diversity in our material world. Chemistry attempts to explain this diversity in behaviour.

While the postlab discussion for experiment 1 brings out the variety in nature, the postlab discussion for experiments 2 and 3 enables the pupils know what chemistry is. No formal definition of chemistry is to be given.

Let the pupils set up the laboratory experiment 4.

Experiment 4 Rusting of iron

Time required : 20 minutes

Ask them to record the observations daily for three days. The purpose of the experiment is to show that rusting is a chemical change. It may also be mentioned that for such a change to occur, the conditions should be conducive. They are the presence of oxygen as well as moisture. Absence of either does not lead to rusting.

Question and Answer

What are the conditions that favour the rusting of iron?

Presence of air and moisture.

Postlab discussion

Emphasise that rusting, a chemical change, occurs only when the interacting materials iron, oxygen and moisture are brought in contact. This fact may be elicited from the pupils after a close observation of what has taken place in the three test tubes. It may be mentioned in passing that rusting is a characteristic tendency of iron to revert to the oxide from which it was originally obtained. Rusting is also a phenomenon which vitally affects the economy of the country and is of concern to technologists and economists.

1.2 Importance of chemistry in daily life

Through a systematic study of materials and the changes they undergo man has learnt to make his life healthier and more comfortable. The material

discussed in the text indicates some areas in which this has been achieved. The following activities for the class are suggested to stimulate the discussion.

- (a) Collection of samples of various fertilizers and recording their uses.
- (b) Collection of pesticides and to list their uses.
- (c) To recollect the preparation of jams and pickles with a view to eliciting the role of preservatives.
- (d) Collection of samples of cloth of natural and artificial fibres, white and coloured.
- (e) Collection of different samples of washing agents like washing soap, washing soda, synthetic detergents and whitening agents.
- (f) Collection of samples of fuels such as coal, wood, kerosene, petrol, spirit (alcohol), etc.
- (g) A visit to a farm of a progressive farmer to study the application of fertilizers.
- (h) A visit to a building construction site with a view to acquaint the pupils with building materials such as mortar, cement, metallic parts, galvanised and tinned sheets, glass plates, etc.
- (i) Drawing diagrams and charts regarding activities mentioned above.

While discussing this topic, the teacher should keep in view the time schedule for this topic. The topic should give the pupils a bird's eye view of the range and scope of chemistry in our daily life. The teacher should emphasise that a specialized knowledge of chemistry is necessary to prepare the things referred to above.

1.3 Matter and the changes it undergoes

Matter

This section describes what we mean by the term matter. It is suggested that no formal definition of matter be given. Present a few materials and point out that they are all examples of matter. See that the materials presented by you include solids, liquids and gases. Impress upon the pupils that by matter we mean anything that occupies space and possesses mass. Pupils readily accept the idea that solids and liquids possess mass and occupy space. It may be difficult for them to extend the same concept to gases. Demonstration I is designed to help them in understanding that air (or any other gas) is also a form of matter.

Demonstration I Air possesses mass

Time required : 15 minutes

Materials required

Uniform light bamboo stick (30 cm. long and about 0.5 cm. in diameter)

Toy balloons of the same weight 2

Cotton thread

Inflator

Procedure

Prepare a toy balance using the bamboo stick. Fill the balloons partially with air and suspend one on each end of the stick and balance them. Puncture one of the balloons with a pin and let the pupils observe that the end carrying the punctured balloon goes up.

While demonstrating the above experiment elicit the following information from the pupils by asking specific questions:

- (i) the punctured balloon weighs less than the inflated one,
- (ii) air possesses mass.

States of matter

Present the various materials representing the three states of matter. Ice, wooden block, marble, paper-weight, water, kerosene, coconut oil, copper sheet, iron powder, salt, air-filled balloon, and such other things may be displayed. The pupils may be asked to mention some more examples.

Demonstration II deals with properties of matter in different states.

Demonstration II Properties of matter

Time required : 25 minutes

Part A Solids

Materials required

Solid objects like paper-weight, brick, marble etc.

Procedure

Show the solid objects and elicit from the pupils that these have fixed shape and volume.

A question may arise here as to why sand, sugar and flour take the shape of the container. This can be explained by directing their attention to the size of individual particles or grains. A single crystal of sugar or sand retains its shape.

Part B Liquids

Materials required

Conical flask (150 ml)

Coloured water

Beaker (150 ml)

Flat bottomed flask (250 ml)

Measuring cylinder (50 ml)

Glass pencil

Procedure

Mark the positions of 100 ml levels on the beaker, conical flask and flat bottomed flask before starting the demonstration.

Measure out 100 ml of coloured water in a measuring cylinder and transfer it successively into the marked containers. Elicit from the pupils that liquids have a fixed volume but variable shape by putting questions such as

- (i) what happened to the volume of the coloured water when it was transferred from one container to another ?
- (ii) what happened to the shape of the coloured water ?

Part C Gases

Materials required

Toy balloons of different shapes	3	Copper turnings or sodium thiosulphate
Watch glasses (5 cm diameter)	3	
Glass vessels of different sizes (beaker, bottle and conical flask)	3	Conc. nitric acid
Dropper		2 ml
Inflator		

Procedure

Pump air into the balloons by giving the same number of strokes in each case. Elicit from the pupils that the same amount of air assumes different shapes.

Place equal amounts (about 1 g) of copper turnings or sodium thiosulphate crystals on each of the watch glasses and add 5 drops of concentrated nitric acid. After adding nitric acid, immediately cover the watch glass with a glass vessel. Let them observe that the brown gas fills each vessel completely. Point out that a gas tends to occupy all available space.

In this part of the experiment, it is necessary to take balloons of different shapes, say cylindrical, spherical and oval. Convince the pupils that the different shapes assumed by air have nothing to do with the amount of air as the same number strokes of the inflator (same amount of air) is used to pump air into each balloon. Glass vessels of any shape but of different sizes must be used.

Demonstration III deals with change of state of matter

Demonstration III Change of state

Time required : 20 minutes

Materials required

Spirit lamp	Ice cubes (pieces)
Hard glass test tube	Freezing mixture (ice+salt)
Stopper with a bent delivery tube	
Test tube	

Procedure

Place 2-3 pieces of ice in a large test tube and warm it on a spirit lamp. Let the pupils observe that ice (solid) changes into water (liquid). Heat further until water boils. Let them note that water (liquid) changes into steam (gas). Close the test tube with the stopper carrying the delivery tube and hold another test tube below the free end of the delivery tube. Draw their attention to the liquid water collecting in the test tube. Place the test tube containing water in the freezing mixture and show them the formation of ice. Point out that the change of state can be reversed by reversing the conditions. Ask the following questions after the demonstration.

1. In how many states can water exist ?

Water can exist in all the three states — solid, liquid and gaseous.

2. How are the changes in state brought about ?

By heating or cooling.

Tell the pupils that some substances do not exist in all the three states. Let them recall the sublimation of ammonium chloride

Types of matter

Mixtures

Most of the materials that we come across are mixtures. In some it is evident, in others it may be necessary to show by experiment. Demonstration IV A is designed to introduce the concept of a mixture to the pupils.

Demonstration IV A Mixtures

Time required : 30 minutes

Materials required

Handlens	Iron filings	6 g
Magnet	Sulphur powder	6 g
Test tube	Carbon disulphide*	10 m
Glass funnel		
Filter stand		
Watch glass (10 cm diameter)		

Procedure

Place some iron filings and sulphur on separate pieces of paper and move a magnet over them. Show that only iron is attracted by a magnet.

* Carbon disulphide is highly inflammable and toxic.
Carbon tetrachloride may also be used.

Place a little sulphur in a test tube and add carbon disulphide to it. Repeat the experiment with iron filings. Let the pupils observe that sulphur dissolves in carbon disulphide whereas iron does not.

Mix 1 g of sulphur powder with 5 g of iron filings on a piece of paper. Let them observe that yellow and black particles are distinctly visible. Move a magnet over this sample and show that only iron particles stick to the magnet leaving sulphur. Shake a little amount of this sample with carbon disulphide. Let the pupils note that only sulphur dissolves leaving iron.

Decant the clear liquid and allow it to evaporate (in cold) on a watch glass. Show that the dissolved sulphur is recovered

Tell them that such a sample is called a mixture. Prepare another sample of iron and sulphur mixture, taking different proportions of iron and sulphur and repeat the experiment.

Let the pupils conclude from the above experiment that

- (i) the composition of a mixture is variable;
- (ii) the components of a mixture retain their properties; and
- (iii) the components can be separated by simple methods like dissolution, filtration and evaporation.

While demonstrating IV A, techniques like dissolution, filtration and evaporation should be shown and explained to the pupils. The handling of bottles and transfer of liquids should be correctly done. The pupils generally imitate their teachers. Hence, due care must be taken even while showing a very simple experiment.

Demonstrations IV B and IV C are designed to allow the pupils to make further study of two more mixtures already familiar to them.

Demonstration IV B Is air a mixture ?

Time required : 30 minutes

Materials required

Evaporating dish	Yellow phosphorus
Trough	
Bell jar with stopper	
Glass pencil	
Iron wire	
Splinter	
Box of matches	
Spirit lamp	

Procedure

General procedure for conducting the experiment is given in the text. The following points should be carefully noted.

- (1) Bell-jar should be graduated previously.
- (2) The mouth of the bell-jar should be immediately closed after the yellow phosphorus is kindled.
- (3) Allow some time for the white fumes to dissolve in water.
- (4) Some phosphorus is left unburnt.

Pupils can be led to the following observations by asking suitable questions.

- (1) The phosphorus burns giving white fumes which dissolve in water.
- (2) The level of water rises by $\frac{1}{5}$ of the initial volume of the air in the bell-jar
- (3) Part of air has been consumed by the burning phosphorus.
- (4) A burning splinter is extinguished by the air left behind in the bell-jar.

Let the pupils conclude that

- (1) a part of air supports combustion. It is the active part of air and is called oxygen,
- (2) the remaining part of air does not support combustion. It is inactive and is nitrogen,
- (3) the air is a mixture mainly of two gases — oxygen and nitrogen.

Demonstration IV C Solutions

Time required : 20 minutes

Materials required

Watch glasses	3	Sea water
Sand bath		Well water
Spirit lamp		Distilled water.

Procedure

Show the pupils bottles marked A, B and C containing samples of sea water, well water and distilled water. Point out that the three samples look alike and are homogeneous. Take 5 ml of sea water, (if sea water is not available, you can prepare it by dissolving 3-4 g of common salt in 100 ml of ordinary water) in a large watch glass and heat it on a sand bath. Show that an appreciable amount of solid is left behind. Place the watch glass containing the residue near the bottle A. Repeat the experiment with well water and distilled water (instead of distilled water you may use rain water collected after the first few showers). Let them observe that well water leaves a little solid whereas distilled water leaves nothing behind.

Which of the three samples is a pure substance? Why?

Distilled water, as it leaves no solid residue.

Tell the pupils that homogeneous materials such as sample A and B containing more than one kind of matter are termed solutions.

Concrete and soil have been cited as other common examples of mixtures. Let the pupils collect soil samples and observe closely. A simple experiment such as sedimentation can easily show that soil is a mixture.

After the demonstration IV C is over, experiment 5 may be given to the pupils on the same day.

Experiment 5 Separation of the components of a mixture

Time required 60 minutes

Proper handling of the apparatus, economic and proper use of reagents and strict adherence to the correct technique should be insisted upon. With the help of any aids, if available, explain the various operations.

Demonstrate to small groups of pupils .

- (i) dissolution of a substance in the minimum amount of solvent,
- (ii) the folding of a filter paper and fitting it in a funnel,
- (iii) the transfer of the solution to the filter ; (solution should be poured only upto the $\frac{2}{3}$ rd, of the filter cone),
- (iv) the evaporation of the solution using sand bath and water bath.

It is desirable to give experiment 5 individually. In the prelab discussion ask the pupils to recall their observations on the solubility of common salt and sand in experiment 1 to bring out the principles governing the choice of solvent. To effect separation, one of the components of a mixture should be soluble in the solvent chosen and the other component should be insoluble or very sparingly soluble.

Compounds

Demonstration V A is intended to show the formation of compound from two simpler substances.

Demonstration V A Formation of sodium chloride

Time required : 15 minutes

Materials required

Pen knife	Sodium metal
Deflagrating spoon	A jar of chlorine
Spirit lamp	
Box of matches	

Procedure

Show the pupils a small piece of sodium freshly cut with a pen knife. Let them note the colour and appearance of the piece. Drop the piece of sodium in a trough of water and show that a vigorous reaction occurs. Show them the bottle containing sodium and explain why sodium is kept under kerosene.

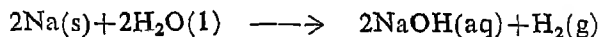
Show them a jar of chlorine. Let them observe the colour. Tell them that chlorine is poisonous.

Warm a little bit of sodium in a deflagrating spoon and introduce it into the jar of chlorine. Let them observe that sodium burns forming a white solid.

Collect the white solid formed and show that it dissolves in water. Ask suitable questions to elicit from the pupils that the white solid has properties different from those of sodium and chlorine from which it is formed.

Tell them that this product is a compound of sodium and chlorine called sodium chloride (common salt).

Sodium causes burns when it comes in contact with skin and so handle it carefully. It floats on water and due to the vigorous reaction it tends to fly out. Therefore, perform this experiment at a safe distance from the pupils and use a small piece (pea size) of sodium for the reaction. The reaction of sodium with water sets free hydrogen and sodium hydroxide solution is formed.



Chlorine can be prepared by heating concentrated hydrochloric acid with manganese dioxide (for details refer to unit 6). Avoid inhaling chlorine as it is poisonous.

It would be proper at this stage, to give a list of compounds with which the pupils are familiar indicating that they can be prepared from simpler substances. Point out that separation of sodium and chlorine from the compound sodium chloride cannot be achieved by simple methods.

In case sodium is not available, the alternative experiment, formation of iron sulphide, may be demonstrated

Alternative to Demonstration V A

Formation of iron sulphide

Materials required

Glass bottle	Iron filings	7 g
Handlens	Sulphur powder	4 g
Test tubes 2	Carbon disulphide	10 ml
	Dilute hydrochloric acid	10 ml

Procedure

Weigh 7 g of iron filings and 4 g of sulphur powder. Mix them well and make it into a paste with minimum amount of water. Put the paste in a glass bottle and press the paste to remove air gaps. Expose the bottle to sun's rays for about 15 minutes. Let the pupils observe the mass turning black and jets of vapour issuing out with a hissing sound. Let them also observe that the bottle becomes hot. After the reaction subsides and the bottle becomes cool remove the black mass on to a piece of paper. Let the pupils contrast the black mass with a mixture of iron and sulphur.

Take a small portion of the black mass into a clean test tube and show that it is insoluble in carbon disulphide (unreacted sulphur present in the black mass dissolves).

Take another small portion of the black mass into a test tube and treat it with dilute hydrochloric acid. Let the pupils observe the evolution of a gas having smell of rotten eggs. This gas is hydrogen sulphide. Show that pure iron filings react with dilute hydrochloric acid to give a colourless and odourless gas (hydrogen).

Let the pupils conclude from the above observations that the black mass has properties different from those of sulphur and iron. Tell them that it is a compound of iron and sulphur called iron sulphide.

If exposure to sun's rays is not possible, a dry mixture of iron filings and sulphur may be heated on a spirit lamp.

In demonstration V A it was hinted that the constituents of a compound cannot be separated by simple methods. However, it is possible to decompose a compound, for example, water, by passage of electric current. This is illustrated by demonstration V B.

Demonstration V B Decomposition of water

Time required . 20 minutes

Materials required

Cup voltmeter	Distilled water	600 ml
Battery or storage cell (6 volts)	Sodium chloride	2 g
Graduated tubes* 2		
Plug key		
Splinter		
Box of matches		

* If graduated tubes are not available select two similar test tubes and mark them.

Procedure

Dissolve a little sodium chloride in 600 ml of distilled water. Use this water to fill a cup voltameter upto half. Fill the graduated tubes also with this water and invert them over the platinum foils (electrodes). There should be no air bubbles in the tubes. Then connect the foils to the terminals of a 6 volt battery. Pass the current for 10 minutes.

The pupils observe that gases are evolved at the two electrodes and that the volumes collected in the two graduated tubes are not same. The volume of the gas collected at the negative electrode is nearly twice that of the one collected at the positive electrode.

Test the gas collected at the negative electrode with a burning splinter. The pupils will notice that the gas burns with a pale blue flame and with a slight explosion. Test the other gas with a glowing splinter and show that the splinter is rekindled.

You may tell the pupils that the gas which burns with a pale blue flame is hydrogen, while the other is oxygen. You may further point out that there are no simple methods to separate hydrogen and oxygen from water.

Repeat the experiment passing current for a longer time. Show to the pupils that the volume ratio of hydrogen to oxygen in water is always 2 : 1.

Thus from the above demonstration, it may be concluded that

- (i) water is a compound of hydrogen and oxygen ;
- (ii) water has properties entirely different from those of hydrogen and oxygen ; and
- (iii) hydrogen and oxygen are present in water in a fixed proportion.

Note : A direct current should be used in this experiment.

It may be mentioned that pure water is a bad conductor of electricity. To make it conducting, sodium chloride is added to water.

Elements

We have seen that water can be decomposed into two new substances, hydrogen and oxygen. It should be brought to the notice of the pupils that there are many other substances which can be decomposed into simpler substances. But substances like hydrogen and oxygen cannot be further decomposed into simpler substances. Such substances are called elements. Some of the elements are listed in the text. Only a few of these elements are common and abundant in nature. The relative abundance of the elements, both free and combined,

in the earth's crust (including surface water and atmosphere) is shown in figure 1.11 in the text. It is apparent that almost one-half of the earth, as we know it, consists of oxygen. Silicon (present in rock and sand) comprises more than one-fourth. In fact, only nine elements make up a total of over 98 per cent of the earth's crust.

Changes in matter

Ask the pupils to recall the changes they observed in experiments 3 and 4. Follow this up by citing some more changes, such as evaporation of water, souring of milk, burning of wood and coal, ripening of fruits, etc

Point out to the pupils that some changes involve only a change in state, for example, melting of wax and evaporation of water. Such changes are called physical changes. In these changes no new substance(s) is formed. On the other hand, in changes like burning of magnesium or decomposition of water by passing electric current, original substance(s) disappears and an entirely new substance(s) is formed. Such changes are called chemical changes.

A chemist is interested in changes because they are responsible for diversity of materials in the world. A systematic study of changes have led to the preparation of thousands of new compounds.

Although we have pointed out two categories of changes in matter, avoid making rigid distinction between the two. No attempt should be made either to enumerate the characteristics of these two types of changes or to give a precise definition of either. There is no sharp demarcation between the two.

In demonstration III reversibility of change of state has been indicated. Although this is generally true of physical changes, it must not be taken as a criterion of physical change.

Conservation of mass

The questions likely to arise now are : 'What happens to matter during a change'? 'Does matter (taken as a whole) have the same mass before and after the change'? Experiments have shown that there is no detectable change in the total mass. This can be illustrated by showing that a flash-bulb weighs the same before and after use. It should be emphasised that for this to be true no material should leave or enter the system under examination, i. e., the system must be a closed one. In an open system there may be loss or gain of mass. For example, lead nitrate on being heated in a test tube loses mass while magnesium gains mass on burning. Demonstration VI is a simple experiment which illustrates the conservation of mass during a chemical reaction.

Demonstration VI Conservation of mass

Time required : 20 minutes

Materials required

Conical flask (250 ml)	Lead nitrate solution (2%)	15 ml
Stopper for the above	Potassium iodide solution (2%)	5 ml
Small test tube		
Thread		
Balance		

Procedure

Take lead nitrate solution in a conical flask. Pour potassium iodide solution in a small test tube which can be placed in the conical flask. Lower the test tube in the flask carefully by means of a thread and hold it in position by fitting the stopper. See that the contents of the test tube do not mix with the solution in the flask. Weigh the flask. Then tilt the flask to mix the two solutions.

Pupils will observe that the colourless solutions of lead nitrate and potassium iodide react to form a yellow precipitate (lead iodide).

Weigh the flask again. Let the pupils note that the weight of the flask remains practically constant. Let them conclude that in a chemical change there is no detectable change in the total mass.

In this experiment any of the following pairs of solutions may be used as reactants.

- (i) Barium nitrate and dilute sulphuric acid
- (ii) Ferric chloride and potassium thiocyanate
- (iii) Lead nitrate and potassium chromate.

BACKGROUND INFORMATION

1.2 Importance of chemistry in daily life

1. Fertilizers

A fertilizer is a natural or manufactured product that promotes the growth of plants and improves the crop yield by supplying adequate amounts of major important elements like nitrogen, phosphorus, potassium, calcium etc. Besides these minor quantities of boron, copper, iron, manganese, molybdenum and zinc are to be supplied. The amount of fertilizer to be applied to the soil depends upon nature of the soil, the average availability of water and the type of the crop.

Natural fertilizers

Farm manures, farm wastes, various seed meals, fish and animal tankages and sewage sludges are widely used. Their contents of nitrogen and other essential nutrients are low. They are largely valued because of their resistance to leaching and their ability to improve texture and biological activity of the soil.

Manufactured fertilizers

(a) Nitrogen fertilizers

Nitrogen contents of some common fertilizers are given below

Ammonium sulphate	...	21.0 %
Ammonium nitrate	..	35.0 %
Ammonium phosphate	..	11.2 %
Sodium nitrate	...	16.5 %
Urea	.	46.6 %

These fertilizers are obtained using atmospheric nitrogen.

(b) Phosphatic fertilizers

Mineral phosphates or phosphate rock, basic slag (a by-product of the steel industry) and bones are the principal sources of phosphorus.

- (i) Super phosphate of lime : It is a mixture mainly of monocalcium phosphate and calcium sulphate. It is obtained by treating tricalcium phosphate with concentrated sulphuric acid.
- (ii) Basic slag : It is a by-product of steel industry and its phosphorus content is 15 to 25%.
- (iii) Bones : These contain about 20% of phosphorus pentoxide.

(c) Potassium fertilizers

Minerals containing potassium chloride and potassium sulphate are the sources of potassium.

2. Pesticides

Any plant, animal or microorganism which is harmful to vegetation is considered as a pest. In our country about 10 million tons of foodstuffs are damaged every year by pests. Chemicals used to destroy pests are called pesticides.

Some common pesticides are

DDT (Dichloro diphenyl trichloroethane)

BHC (Benzene hexachloride)

ANTU (Alpha naphthyl urea)

Bordeaux mixture (Copper sulphate + hydrated lime + water)

Mercuric chloride

Sodium trichloroacetate

3. Germicides, antiseptics and disinfectants

A germicide is any chemical that kills microorganisms that cause disease. An antiseptic is a substance that arrests the activity, halts the growth, or otherwise renders harmless most of the harmful microorganisms without being injurious to human tissue. A disinfectant is a substance that halts the growth of and destroys microorganisms.

4. Food preservation

Food decays due to the action of microorganisms and enzymes. In order to preserve it, methods such as drying, smoking and salting are used. Removal of moisture retards the growth of microorganisms. Higher concentrations of salt or sugar also arrest their growth. Smoking slows the action of microorganisms and enzymes and imparts a distinctive flavour to the food. Canning, refrigeration, use of chemical additives and radiation are the methods commonly used today.

5. Fibres, plastics and rubber

These are the materials which are built up from very large molecules. The different arrangements of these molecules decide whether the substance is suitable for obtaining flexible fibres, hard plastics or elastic rubbers. Molecules made up of a large number of repeating units are called polymers. The process of polymerisation is responsible for a variety of synthetic fibres, plastics and rubber. Rayon is an artificial silk, but it differs from the latter because it is made from cellulose and not protein. Nylon is also a synthetic fibre.

1.3 Matter and the changes it undergoes

1. Change of state

In the text reference is made to the changes in state of tin and iron. The relevant data are given below :

	Melting point	Boiling point
Tin	232 °C	2687°C
Iron	1535° C	2885°C

Sublimation

The process of sublimation involves spontaneous conversion of a solid substance into its vapour when it is heated below its melting point. If the temperature goes above the melting point it will liquefy and then vapourise like other substances.

2. Composition of air

In the text reference is made to the fact that besides nitrogen and oxygen, there are other gases in air. The presence of carbon dioxide in air can be demonstrated by blowing air into lime water for about ten minutes. The lime water turns milky establishing the presence of carbon dioxide in air. Instead of blowing air into lime water, the pupils may be asked to expose lime water to air for a day or two and observe the crust that is formed on the surface.

Lime water can be prepared by shaking 2 to 3 g of slaked lime in one litre of water. Let it stand for two or three hours and then decant. Store the clear solution in a stoppered bottle to protect it from atmospheric carbon dioxide.

To establish the presence of water vapour in air, a few ice pieces can be placed in a beaker and the mouth of the beaker closed with a watch glass. After a while, the formation of droplets of water outside the beaker will be observed. Since the droplets could not have come from ice, they must have come from air around the beaker. The formation of clouds, fog and dew may be cited as other examples to establish the presence of water vapour in air. Besides the above, rare gases are present in air in small quantities. They are chemically inactive and are also called inert gases. However, under special circumstances they may form compounds with other elements.

Average volume composition of dry air

Nitrogen	...	78.03%
Oxygen	..	20.99%
Argon	..	0.94%
Carbon dioxide	..	0.03%
Other rare gases	...	0.01%

The composition of air varies slightly from place to place.

Quiz

1. Fill up the blanks .
 - a. Water is a of hydrogen and oxygen.
 - b. A solid can be changed into a liquid by
 - c. Burning of charcoal is a ... change.
2. Tick off the correct answer :
 - i. Sea water is
(a) a mixture (b) a compound (c) an element
 - ii. When a chemical reaction takes place there is
(a) an increase in the total mass
(b) decrease in the total mass
(c) no change in the total mass
3. Why is copper called an element ?
4. 1 g of a substance A on being heated in air gave 1.5 g of a substance B. Substance B was not decomposed by further heating. Could the substance B be an element ?
5. Why is magnesium oxide classified as a compound ?
6. A white powder when heated strongly left a yellow solid residue and gave off a gas which turned lime water milky. Was the white powder an element or a compound ?
7. What are the characteristics of a chemical change ?

Structure of Matter

INSTRUCTION MATERIAL

Introduction

The pupils are made familiar with some substances in unit 1. They studied the properties of some of them and found that all substances are not alike. They will therefore, be curious to know what matter is made up of and why one kind of matter is different from another.

To find answers to such questions they are led to study the behaviour of gases. Simple experiments devised enable them to understand the particle (molecular) nature of matter and to explain the differences in properties of solids, liquids and gases.

A molecule is made up of smaller particles called atoms. An elementary idea of the sub-atomic nature of matter is also given. This is helpful in understanding ionic valence and for writing the formulae of ionic compounds. Further it helps in understanding the properties of solutions dealt with in latter units.

Outline

- | | |
|--|--|
| 1. Behaviour of gases (2.1). | 5. Modern picture of the atom (2.5). |
| 2. Nature of solids and liquids (2.2). | 6. How heavy are atoms and molecules? (2.6). |
| 3. Change of state (2.3) | 7. Molecular weight (2.7). |
| 4. Atoms and molecules (2.4). | |

Concepts

1. Matter is made up of particles called molecules.
2. Molecules consist of atoms of the same or of different kinds.
3. Elements are substances containing atoms of the same kind ; compounds contain atoms of different kinds.
4. An atom is made up of electrons, protons and neutrons.
5. Atomic and molecular weights are relative weights of atoms and molecules respectively.
6. Gram-atom and gram-molecule are the weights of 6×10^{23} atoms and molecules respectively. This number is known as 'Avogadro's number'.

SCHEDULE

Lesson	Experiment	Postlab	Demonstration	Text covered
1	—	—	I and II	—
2	6 and 7	Discussion	III	—
3	8	Discussion	IV	2.1
4	—	—	—	2.2 to 2.4
5	9	Discussion	V	2.5
6	—	—	—	2.6 and 2.7
7	—	—	—	Review and quiz

Development

2.1 Behaviour of gases

One of the important properties of a gas is diffusion — a tendency to move from place to place. This idea is not to be put forward to the pupils as a definition. Discussion on this phenomenon may be initiated by asking them how they perceive the smell of spices or roasted food from the kitchen, the fragrance of flowers or the pleasant smell when a bottle of perfume is opened even when they are away from the source of the smell. Then demonstrate the diffusion of acetic acid vapour and nitrogen dioxide gas.

Demonstration I Diffusion of gases

Time required : 30 minutes

a. Diffusion of acetic acid vapour

Materials required

Watch glass (10 cm diameter)

Glacial acetic acid* 2 ml

*Liquor ammonia or perfume may also be used.

Procedure

Close the doors and windows of the room and place a watch glass in a corner. Add 2 ml of glacial acetic acid to the watch glass. Ask the pupils to raise their hands as soon as they perceive any smell. (Pupils nearest the watch glass containing the acid will first raise their hands and those farthest from the watch glass will raise their hands last.) Draw their attention to this aspect and explain that the vapour of the acid takes time to move from the source to all parts of the room.

b. Diffusion of nitrogen dioxide

Materials required

Gas jars	2	Sodium thiosulphate	1 g
Glass plate (10 cm square)		Conc. nitric acid	5 ml

Procedure

Place 5 ml of concentrated nitric acid in a gas jar. Add about 4 or 5 crystals of sodium thiosulphate to the acid. A brown gas, nitrogen dioxide, is evolved. Cover the jar with a glass plate and wait till the gas fills the jar completely. Remove the plate and invert another gas jar over the first. Have the pupils observe the diffusion of the gas into the upper jar. (In a short time the gas fills both the jars). Elicit from the pupils that gases move from place to place and occupy the entire space available.

Explain to the pupils the phenomenon of diffusion on the basis of the particle nature of gases.

Ask the pupils whether the particles of different gases are of the same mass. Accept any answer given by them; let them observe the relative rates of diffusion of ammonia and hydrogen chloride to check their answer.

Demonstration II Relative rates of diffusion of gases

Time required : 15 minutes

Materials required

Glass tube* (30 cm long 2.5 cm diameter open at both ends)		Conc. hydrochloric acid	2 ml
Corks	2	Liquor ammonia	2 ml
Iron stand			
Metre scale			
Cotton plugs	2		
Droppers	2		

(*Clean and dry the glass tube well before the demonstration).

Procedure

Insert two cotton plugs into the glass tube, one at each end. Fix the tube horizontally to a stand. Moisten simultaneously the two plugs, one with

concentrated hydrochloric acid and the other with liquor ammonia using separate droppers. Close the two ends of the tube with corks.

Let the pupils observe the formation of a white cloud in the form of a ring within the tube. (This is due to the formation of ammonium chloride as a result of reaction between ammonia and hydrogen chloride diffusing from either end of the tube). Show the action of HCl on NH_3 . Measure the distance of the ring from each end of the tube. Elicit from the pupils which gas moves faster. Ask the pupils to calculate the ratio of the speeds.

Note :— The speeds of diffusion are roughly proportional to the distances covered by the two gases.

Tell the pupils that the speed of a gas depends on the mass of the particles ; the heavier the particles of the gas, the lesser is its speed. Help them to infer that the particles of hydrogen chloride are heavier than those of ammonia.

You may then deal with the other property of gas – ‘Gas exerts Pressure’ Expt. 6 gives the pupils an idea that air in a balloon exerts pressure and so the walls are under tension. The tension increases as more and more air is blown in.

Experiment 6 Gas exerts pressure

Time required : 15 minutes

Postlab discussion

In the postlab discussion explain gas pressure. The particle nature of matter may be used for this also. The pressure exerted by a gas is due to the particles striking the walls of the container. Ask the pupils why the balloon bulges when more and more air is blown into it. The balloon also bulges nearly uniformly on all sides.

Explain all these observations on the basis of particle nature. The particles are uniformly distributed ; so they hit the sides of the balloon equally on all sides bulging the balloon uniformly.

The next question that arises is whether the particles in a gas are close to one another or there are spaces between them. Answer to this may be found in a study of the pressure-volume relationship of a gas.

Experiment 7 Effect of pressure on the volume of a gas

Time required : 15 minutes

Mention that the temperature of the gas is to be kept constant when the pressure-volume relationship is considered.

Questions and answers

When the piston is pushed half way into the barrel

- a. what happens to the volume of the air in the barrel?
 - b. does the pressure of air in the barrel increase or decrease?
-
- a. The volume of the air decreases.
 - b. The pressure increases.

Postlab discussion

A certain quantity of air is enclosed in the barrel of the piston. The piston is pulled back and the outlet closed. When the piston is pushed forward the same amount of air now occupies less volume. Pushing the piston forward increases the pressure on the gas. When the piston is pulled back the volume of the air increases. Pulling it back decreases the pressure. Therefore, as the pressure increases the volume of air in the barrel decreases and vice-versa. The same idea is illustrated by the following demonstration.

Demonstration III Effect of pressure on the volume of a gas

Time required : 10 minutes

Materials required

Glass tube (50 cm \times 0.6 cm diameter)

Rubber tubings (2 cm long 0.5 cm diameter) 2

Cycle valve tube

Glass rod (1 cm long 0.7 - 0.8 cm diameter)

Syringe (5 or 10 ml)

Set up the apparatus as shown in figure well before the demonstration.

Bend a 25 cm long tube as shown in figure 2.1. Add coloured water into the long arm of the tube till the height of the water is about 12 cm in both the arms of the tube. Attach a 2 cm rubber tubing to the bent tube. Slip a piece of cycle valve

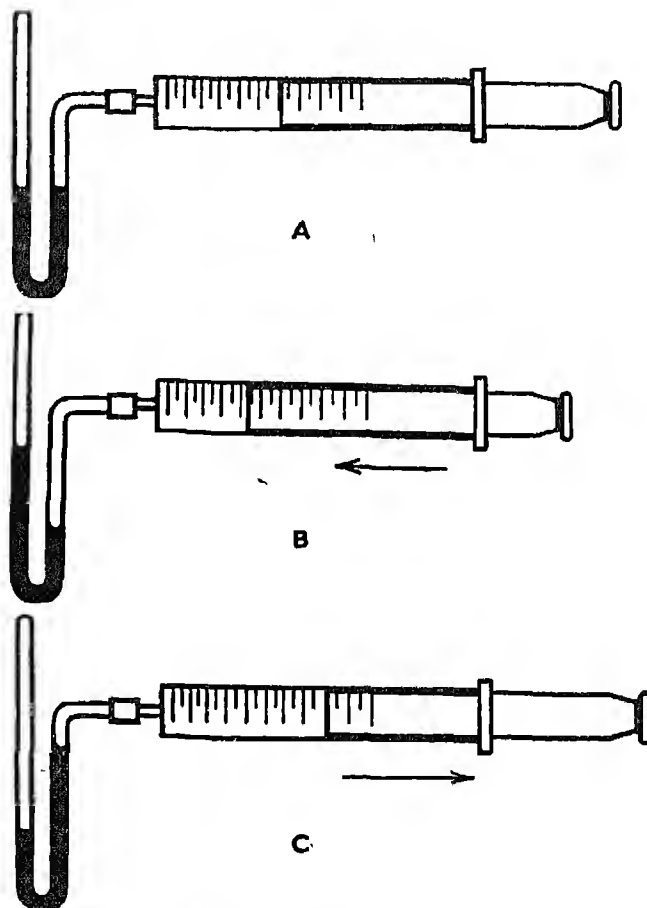


Figure 2.1 Effect of pressure on the volume of a gas

tubing over the nozzle of the syringe and pull the piston to the middle of the barrel. Insert the nozzle of the syringe into the rubber tubing attached to the small arm of the tube. Close the other end of the tube with another rubber tubing and a glass rod. Equalise the level of the liquid in both the limbs by slightly adjusting the position of the piston. Tie a scale to the tubes or paste a graph sheet. The apparatus is ready for use. A definite amount of air is enclosed in the long arm.

Procedure

Note the length of the air column in the long arm of the tube and the position of the piston (A in figure). Push the piston into the barrel by about two divisions on the scale marked on the barrel and hold the piston in that position (B). Let the pupils observe the changes in the height of the water level in the two arms. Note the level of the water in the long arm. Allow the piston to go back to its original position and let the pupils observe the water level in both the arms.

Pull the piston back by about two divisions (C) and let the pupils note the levels of water in the long arm.

Explain that when the pressure is increased by pushing the piston in, the volume of the air enclosed in the left arm decreases. It increases when the piston is drawn out and the pressure decreased.

The decrease in volume with increasing pressure reveals that there are empty spaces between molecules in a gas. They are not close to one another. When pressure on the gas is increased the molecules are brought closer. Thus compression occurs.

The effect of temperature on the volume of a gas can be studied by experiment 8.

Experiment 8 Effect of temperature on the volume of a gas.

Time required: 30 minutes

Precautions

See that the pupils use a dry boiling tube. The tube should be heated till about 10 bubbles come out. During cooling, if the water level rises up too high, ask the pupils to remove the delivery tube from the beaker. Otherwise water enters the hot boiling tube and causes breaking.

Questions and answers

1. What happens to the volume of air when its temperature is
a) decreased and b) increased?

The volume of air increases when the temperature is increased and it decreases when the temperature is decreased.

2. Explain why cycle tubes sometimes burst in summer

When the tyres are on a hot road the air enclosed in the tubes expands and exerts high pressure resulting in bursting of the tubes.

The volume-temperature relationship may also be illustrated by the following demonstration.

Demonstration IV Effect of temperature on the volume of a gas

Time required 20 minutes

Materials required

Boiling tube

Capillary tube (50 cm long 1 mm diameter)

Glass tube (10 cm) bent at 90°

Glass rod 1 cm x 0.4 cm

Rubber tubing 3 cm

Rubber cork (2 holed)

Iron stand

Spirit lamp

Mercury 2 ml

Get the apparatus (figure 2.2) ready before starting the demonstration.

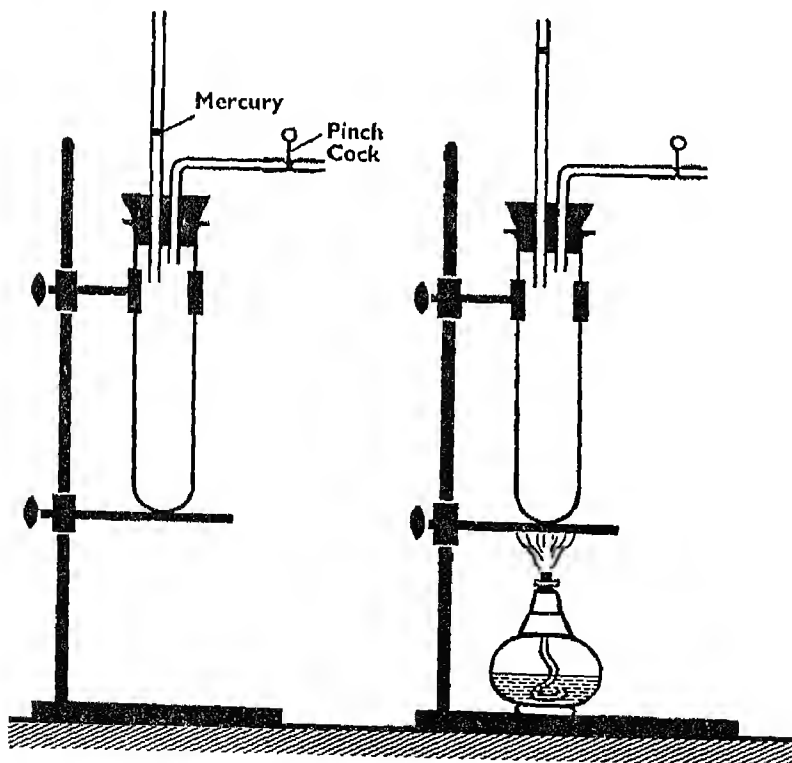


Figure 2.2 Effect of temperature on the volume of a gas

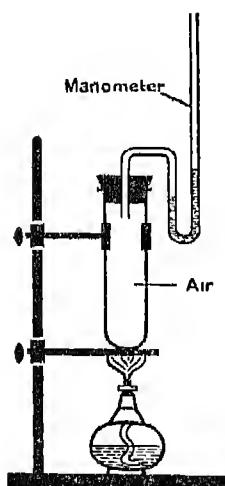
Fit the boiling tube with the rubber cork. Insert the bent tube into it through one of the holes. Introduce into the capillary a small amount of mercury to act as a seal. This can be done as follows: Dip one end of a warm capillary tube in a pool of mercury, close the other end tight with the fore finger. Mercury enters the capillary as it cools. Allow a very small amount of mercury to enter the tube. Turn it horizontal. Tap it gently so that the mercury seal moves to the middle of the tube. Introduce the capillary tube into the other hole of the rubber stopper. Close the bent tube with a piece of rubber tubing and a piece of glass rod. See that the cork is air-tight (apply wax if necessary).

Procedure

Mount the boiling tube vertically as shown in figure 2.2. Mark the position of the mercury seal. Place a spirit lamp 10 cm below the bottom of the boiling tube. Let the pupils observe the position of the mercury seal as the tube is warmed. Stop heating when it starts moving up.

Discussion

The volume of a gas increases with increase in temperature. On heating, the molecules of a gas pick up more energy and move apart. The distances between molecules increase. The volume occupied by a gas, therefore, increases.



In case mercury is not available the experiment may be demonstrated in the following way. The apparatus used may be made as follows.

Fit a boiling tube with one holed rubber stopper. Through the hole insert a glass tubing bent at right angles. Take a manometer tube (the one used in Demonstration III) and fill it with some coloured water such that the levels of the liquid in the two limbs are about 12 cm high. Connect the bent tube and the manometer tube by means of a rubber tubing. Note the level of water in the longer limb. Heat the boiling tube with a spirit lamp for a couple of minutes and note the level of water again. Explain that the rise in level is due to the expansion of air in the tube.

Figure 2.3 *Effect of temperature on the volume of a gas*

2.2 Nature of solids & liquids

Recall to the pupils the change of state of matter studied in unit 1. A gas on cooling gives a liquid which on further cooling gives a solid. Remind the pupils of the assumption that gas contains particles. Let them guess what liquids and solids contain. Elicit that they too contain particles.

Then focus their attention to the differences in the properties of solids, liquids and gases. Ask why they differ in their properties though they are made up of the

same particles. These differences could be explained in terms of intermolecular distances and attractive forces. Draw the attention of the pupils to the volumes of the same mass of gas, liquid and solid. Elicit what would be the distances between particles in the three states of matter.

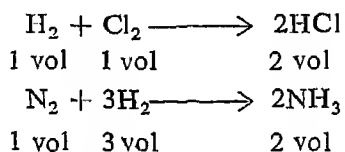
Making use of the differences in the distances between molecules in the three states, you may bring in the differences in forces holding the molecules in the three states. The forces are strong in solids, less so in liquids and weak in gases. These differences in distances and attractive forces between molecules may be used for explaining why solids have definite volume and shape where as liquids have definite volume but no definite shape and gases have neither a definite volume nor shape.

Change of state of matter

Recall to the pupils the need for supply of energy to convert a solid to a liquid and liquid to vapour and the energy release during the reversal of the process. Explain this in terms of molecular attractions.

2.3 Is a molecule divisible?

Now that the pupils know about the molecular nature of matter, explain that a molecule is not the simplest particle but made up of atoms by recalling the demonstration on electrolysis and other examples of combining volumes of gases.



Making use of Avogadro's hypothesis deduce the divisibility of molecules into atoms.

It may, however, be mentioned here that atomic nature of matter was first proposed by John Dalton to explain the laws of chemical combination. (Laws of conservation of mass and definite proportions etc.) Cataloguing postulates of atomic theory may be avoided. The postulates, on the other hand, may be used for explaining the formation of molecules.

Modern picture of the atom

The pupils should not go away with the impression that an atom is the smallest particle of matter. It is necessary to give some elementary idea that an atom contains still smaller particles. Before you talk about these sub-atomic particles, it is necessary to impress upon the pupils that matter is electrical in nature.

An experiment on the production of static electricity may serve the purpose to some extent. Let the pupils do experiment 9

Experiment 9 Electrical nature of matter

Time required : 10 minutes

Question and answer

Did you ever notice hair on your skin standing up on taking off a terylene shirt? If so, give the reason for it

Yes. The shirt constantly rubs the hair on the skin and both get electrically charged. When the shirt is removed it attracts the hair, which consequently stands up.

The electrical nature of water can also be illustrated by the following demonstration.

Demonstration V Electrical nature of matter

Time required . 15 minutes

Materials required

Ebonite rod^{*}
Terylene
Burette** 50 ml
Burette stand
Beaker 400 ml

Procedure

Clamp a burette to a stand and fill it with water. Open the stop cock and allow a thin and steady stream of water to run down. Let the pupils observe the direction of the stream. Bring one end of an ebonite rod near the stream as close as possible without touching the water. See that the rod is not much below the tip of the burette nozzle. Let the pupils observe the direction of flow of water.

Rub the rod briskly with a piece of terylene about 20 times and bring it near the stream of water as above. Let the pupils observe the change in the direction of flow.

Discuss how the ebonite rod acquires the ability to attract water. When it is rubbed with terylene, it acquires an electric charge. When it is brought near water, it attracts water. It is evident from this that the ebonite rod and water contain charged particles in them. Similar is the case with other materials.

With this introduction about the electrical nature of matter, you may then introduce the idea that the atom, a neutral particle of matter contains negatively charged particles (electrons), positively charged particles (protons) and particles without charge (neutrons). The neutrons and protons are concentrated in a small space at the centre of the atom called nucleus. Electrons are distributed around the nucleus. It is possible to remove some electrons from an atom easily and

(*Plastic comb may also be used; make it dry before use)

**Any arrangement that gives a thin stream of water may be used

transfer these to other atoms. Atoms which have lost or gained electrons are called ions. Those which have lost electrons are positively charged and are called positive ions and those which have gained electrons are negatively charged and are called negative ions.

Atomic weights

Impress upon the pupils that actual weight of an atom is very small as it is a very tiny particle. Its weight is so small that it cannot be conveniently used in chemical calculations. Weights of atoms are compared with an atom of one of the elements chosen as standard. Discuss the choice of hydrogen as standard and its limitations. Mention that its place is now taken by carbon taken as 12. Pupils should know that atomic weight is not the actual weight of the atom but is only a relative weight. Also tell them that when expressed in grams it is the weight of 6×10^{23} atoms.

Molecular weights

Molecules are also very small. It is difficult to use their actual weights. So only comparative weights are used. The weight of a molecule is compared with the weight of one atom of carbon taken as 12. It is the weight of 6×10^{23} molecules of the substance.

Since molecules are made up of atoms, molecular weights may be taken as equal to the sum of the weights of the atoms present in the molecule.

BACKGROUND INFORMATION

2.2 Differences in the properties of solids, liquids and gases

In the gaseous state we know that the particles are far apart. The mean free space between the two particles is several thousand times their dimensions. In addition the particles are in constant random motion at high speeds causing a very large number of collisions among themselves and with the walls of the vessel. There is thus a very high degree of disorder in the distribution of the particles in space. On the other hand in liquids, the particles come nearer and some order begins to set in. In this case the particles are held together by weak attractive forces. Energy in the form of heat has to be supplied to convert a liquid into its vapour. Thus a liquid has less energy content as compared to that of a gas. In solids (non-ionic) the particles are held together by forces of intermolecular attraction and they are arranged in an orderly way i.e. randomness is at a minimum. This is also a state of minimum energy. In ionic solids the oppositely charged ions are held together by stronger electrostatic attractions and that is why they have high melting points.

A crystalline state may be compared with that of a class room in which the seats are arranged regularly and boys and girls are seated in a definite order, say, alternately. Thus liquid state may be compared to the same class room where they sit at random and a gaseous state may be compared to a class room where there are no seats and boys and girls are running about in all directions and there is no order whatsoever.

Thus in the gaseous state there is maximum disorder (randomness), in liquids there is some order (less randomness) while in solids there is maximum order (least randomness).

The process of change of state is one of absorption or liberation of energy.

2.4 Early ideas about matter

The ancient Hindu philosophers visualized matter as consisting of ultimate particles. During the 7th and 6th century B. C., the Sankhya and the Nyaya Vaiseshika system of philosophy advanced the ideas of dual forces as the origin of the entire universe and the theory of five elements constituting all matter. These two forces were named as Purush and Prakriti and the five elements as Kshiti (earth), Ap (water), Tej (fire), Marut (air) and Vyom (space).

The idea that all matter is made up of minute indivisible particles, called atoms, also originated with the ancient Hindu and Greek Philosophers. Kanada, the founder of Vaiseshika system of Indian philosophy propounded the atomic theory of matter (6th century B. C.).

In the Greck period, there were two schools of thought. One school, led by Aristotle believed that an iron bar, for example, could be sub-divided into two pieces and further sub-divided into four, eight, sixteen, etc pieces. Aristotle claimed that there is no limit to the number of times a piece of substance can be subdivided, an infinite number of sub-divisions could be made without changing the chemical composition of the substance. Democritus, however, led the school that believed that although a piece of iron can be sub-divided many times, eventually a point is reached where it is impossible to further sub-divide the substance and still retain the characteristic properties of that substance. These small particles are now called atoms.

2.5 Electrical nature of matter

Water is electrically neutral. However, in water concentration of electrons around oxygen is more as compared to that around hydrogen i.e. in a neutral water molecule the hydrogen end is positive and the oxygen end, negative. This separation of charges in a molecule without breaking a bond leads to what is called a 'dipole' and such a molecule is said to be polar.

On bringing a charged comb near a stream of water the polar water molecules are attracted, thus bending the stream of water from its straight path.

STRUCTURE OF ATOM

For quite a long time, the atom was considered to be the smallest particle of matter. Researches during the late 19th and the present century showed that the atom is not the smallest particle of matter, but it contains still smaller particles. After the discovery of voltaic cell (a source of electric current) attempts were made to pass electric current through solids and liquids to see if there are any changes brought about. When encouraging results were obtained on passing electric current through solutions, scientists began attempts to pass it through gases. The apparatus employed consisted of a closed glass tube with electrodes sealed in it at each end. The tube is connected to a vacuum pump, so that it could be partially evacuated (Figure 2.4).

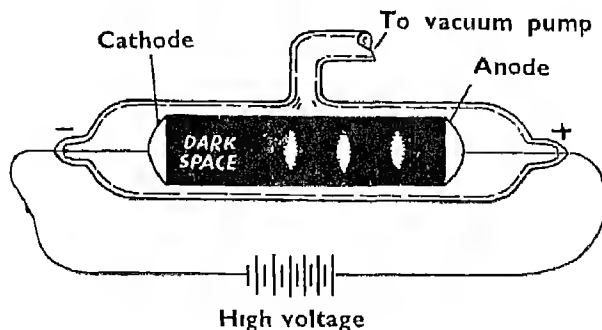


Figure 2.4 Discharge of electricity through gases

When a high potential of the order of 10-20 kilo volts was applied to a gas at normal atmospheric pressure, the gas did not allow current to pass through, but did allow it when the pressure was reduced to 1.0 mm. A glow appeared when the pressure was reduced to 0.1 mm. On reducing the pressure to 0.001 mm the glow disappeared but the glass wall opposite the cathode, the negative electrode, became fluorescent with a green light. This was supposed to be due to some sort of radiations striking the walls. These radiations came from the cathode and so they were given the name cathode rays.

A study of the properties of these cathode rays revealed the existence of a new particle. These rays cast a well-defined shadow of an object placed in their path showing that the rays travel in straight lines as ordinary light rays. Are the rays same as the light rays? To examine this, a small paddle wheel was placed in the path of the rays. The paddle wheel started rotating. This can happen only when the blades of the paddle wheel are struck by some material particles moving with high speeds. Ordinary light rays are not capable of rotating the wheel. It

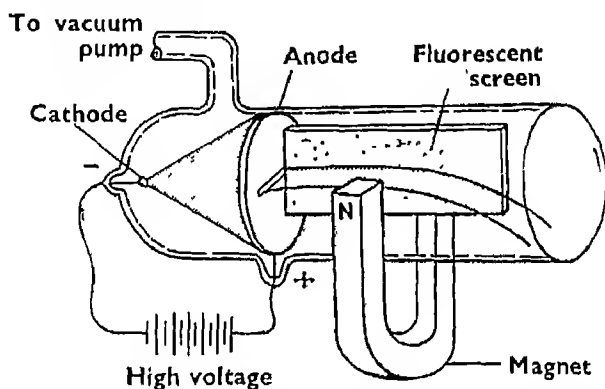


Figure 2.5 *Effect of magnetic field on cathode rays*

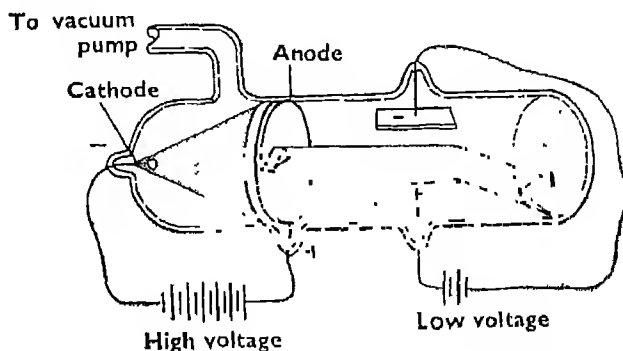


Figure 2.6 *Effect of electric field on cathode rays*

was, therefore, believed that the cathode rays contain minute particles. These rays were deflected by a magnet showing that the particles of the rays bear an

electric charge. They are also deflected towards the positive pole when an electric field is applied. To see what type of charge is borne by these particles the rays were passed into a metal box connected to an electrometer (a device to show the nature of the charge). The electrometer indicated that the rays bear a negative charge. The particles present in the cathode rays are very minute and bear a negative charge. These particles were given the name electrons by Perrin. Many gases gave these electrons when they were subjected to a strong electric discharge. *The point of interest is that the electrons obtained from different gases were found to have the same properties.* These electrons are also obtained from heated metal filaments or by exposing the metals to ultraviolet rays. *Electrons obtained from these sources also show the same properties. They are identical. Electrons are a common constituent of all matter.*

If electrons are obtained by the disintegration of the atom, the latter should have some positive charges in it to neutralise the charge of electrons because the atom is electrically neutral. This points out to the possibility of the existence of positively charged particles in atoms. If this be true, the gas atoms robbed of a few electrons (negatively charged particles) in the discharge tube would naturally acquire a positive charge. These positive residues will move away from the anode (positive electrode) and dash against cathode (negative electrode). If a hole is cut in the cathode, these particles would pass through the hole and appear behind the cathode as a stream. Such rays were noticed when a perforated cathode was used. These rays were thought to come from the anode and so these were called anode rays.

The properties of these rays were studied by Sir J. J. Thomson. They were deflected by electric and magnetic fields just as cathode rays, but in the opposite direction, showing that unlike cathode rays they consist of positively charged particles. Masses of these particles were found to be the same as those of the atoms from which they were derived and found to be equal to the atomic weight of the gas in the tube. Thus while the electrons are of negligible mass the positive particles have nearly the whole mass of the atoms.

The study of the anode and cathode rays obtained as a result of the breaking down of gas atoms in the discharge tube, shows clearly that the atom is made up of negative electrons and positive particles. Where are these positive and negative particles situated in the atom?

Lord Rutherford was engaged in the study of the scattering of α -particles (Helium atoms from which the electrons are removed). When they were made to bombard the atoms in a metal at high speed, many of them went straight through the metal. A few of the α -particles suffered large deflection and very few of them were rebounded. Why are only a few deflected and not many? α -particles are helium atoms carrying two positive charges, they could be deflected only by something positively charged. To account for these observations Rutherford assumed that positive charge in an atom is concentrated in a small space at the centre. This is called the nucleus. Thus there is a large empty space in the atom. The α -particles which went closer to the nucleus suffered a deflection. Because of large empty spaces in the atom,

most of the α - particles pass through the empty space of the atom and are unaffected by the nucleus as they are far from it. Only a few α - particles go closer and get repelled by the nucleus (Figure 2.7).

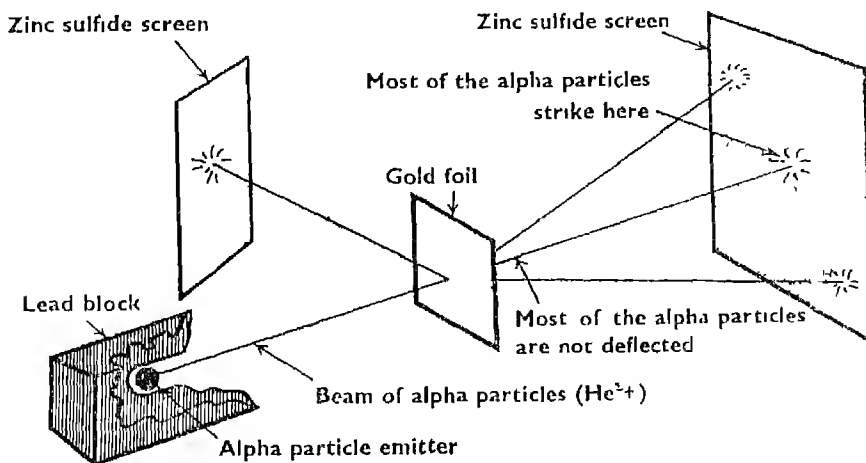


Figure 2.7 Schematic diagram of Rutherford's scattering experiment

Since the atom is electrically neutral, the negative electrons must be distributed in space around the nucleus to balance its positive charge.

Quiz

- One million oxygen molecules react with sufficient hydrogen to form water molecules. How many water molecules are formed? How many hydrogen molecules are consumed?
- Sulphur dioxide gas has a molecular weight of 64; chlorine gas has a molecular weight of 71. Suppose a gas-jar of chlorine was placed on top of a gas-jar of sulphur dioxide with the open ends of the jars together. Which of the following would you expect to have happened after a few minutes?
 - The two gases would not have mixed at all.
 - Some of the Chlorine would have moved into the sulphur dioxide jar, but some of the sulphur dioxide would have moved into the chlorine jar.
 - Some of the sulphur dioxide would have moved into the chlorine jar, but none of the chlorine would have moved into the sulphur dioxide jar.

3. Match the expression in the left hand column with an expression in the right hand column which best completes a correct statement regarding the molecular nature of matter.

- | | |
|--|---|
| (i) Gas molecules move with | a. contain equal number of gas molecules |
| (ii) Equal volumes of gases at the same temperature and pressure | b. constant average speed in different directions |
| (iii) The space is filled by the gas because of | c. with increase in temperature |
| (iv) The volume occupied by a given mass of gas increases | d. the rapid motion of the gas molecules |
| (v) The volume occupied by a given mass of gas decreases | e. with increase in pressure |

4. Fill up the blanks choosing the correct words.

- To overcome the attractive forces between molecules energy is (absorbed, released)
- On heating a gas the molecules in it move. (nearer to each other, apart)
- On cooling a gas, its molecules (lose energy, gain energy)
- Gram atomic weight of an element is the weight of (one atom, 6×10^{23} atoms)
- Molecules of a gas move (slower, faster) at a higher temperature.

5. State whether the following statements are True or False.

- Neutron is a particle carrying a negative charge.
- The volume of a given mass of gas increases with increase in temperature.
- Gases diffuse because of empty spaces between the molecules and the tendency of molecules to be in motion.
- Pressure of a gas increases with decrease in temperature.
- 2 volumes of hydrogen and 1 volume of oxygen contain the same number of molecules at the same temperature and pressure.
- Molecules in a gas are closer to each other than in a solid.

Symbols, Formulae & Equations

INSTRUCTION MATERIAL

Introduction

The purpose of this unit is to introduce the pupils to the language of chemistry. The use of symbols, formulae and equations to represent elements, compounds and chemical reactions are explained in the text.

Outline

1. Symbols of elements (3.1).
2. Meaning of a symbol (3.2).
3. Formulae of molecules (3.3).
4. Valency and chemical formulae (3.4).
5. Chemical equations (3.5).

Concepts

1. A symbol is a short and convenient form of representation of an element.
2. A formula represents the ratio of atoms of the various elements present in a substance.
3. Valency is the combining capacity of an element.
4. An equation is a short and simple representation of a chemical reaction.

SCHEDULE

Lesson	Experiment	Postlab	Demonstration	Text covered
1	—	—	—	3.1 & 3.2
2	—	—	—	3.3
3	—	—	—	3.4
4	—	—	—	3.5
5	—	—	—	Review and quiz

Development

3.1 Symbols of elements

A good starting point for the study of symbols is to quote some examples from everyday life, such as the use of road signs and initials of name to impress on the pupils the convenience of using symbols. Mention may also be made of the use of mathematical symbols. This can be followed by a discussion of the method of representing chemical elements by symbols.

There are 103 chemical elements now known and each has been assigned a particular symbol.

3.2 Meaning of a symbol

You may explain the meaning of symbol by taking some examples other than those given in the text like O and N. The symbol O or N represents one atom or one gram atom of oxygen or nitrogen depending on the context. One gram atom of oxygen (16 g) or nitrogen (14 g) contains the same number, 6×10^{23} , of atoms of oxygen or nitrogen.

3.3 Formulae of molecules

While dealing with the formulae of molecules, first take up examples of molecules containing only one type of atoms e. g. O_2 , N_2 , Cl_2 etc. Then take up examples of molecules which contain different types of atoms, e. g., HCl , CO_2 , NH_3 , etc.

Molecules may be mono-, di- or poly atomic, i. e., their molecules may consist of one, two or more atoms.

Molecules	Examples
Monatomic	Helium (He), neon (Ne), argon (A) various metals, in vapour state, such as Hg, Na, etc.
Diatomic	Hydrogen (H_2), oxygen (O_2), nitrogen (N_2) etc.
Triatomic	Ozone(O_3), carbon dioxide(CO_2).
Polyatomic	Sulphur (S_8), ammonia (NH_3), methane (CH_4).

Calculation of percentage composition from formula

The pupils should be given sufficient practice in calculating the percentage composition from a formula using the table of atomic weights. The knowledge of the percentage composition of a compound is useful in analytical and industrial operations.

The calculation of the percentage composition of potassium nitrate may be given as an additional exercise. The details of the calculation are given below. The formula of potassium nitrate is KNO_3 .

$$\begin{aligned}\text{Formula weight} &= \text{K} + \text{N} + \text{O} + \text{O} + \text{O} \\ &= 39 + 14 + 16 + 16 + 16 = 101\end{aligned}$$

Percentage of potassium

101 g of KNO_3 contain 39 g of potassium

$$\therefore 100 \text{ g of } \text{KNO}_3 \text{ will contain } \frac{39 \times 100}{101} = 38.6 \text{ g of K}$$

Hence the percentage of potassium in $\text{KNO}_3 = 38.6$

Percentage of nitrogen

101 g of KNO_3 contain 14 g of nitrogen

$$\therefore 100 \text{ g of } \text{KNO}_3 \text{ will contain } \frac{14 \times 100}{101} = 13.9 \text{ g of N}$$

Hence the percentage of nitrogen in $\text{KNO}_3 = 13.9$

Percentage of oxygen

101 g of KNO_3 contain 48 g of oxygen

$$\therefore 100 \text{ g of } \text{KNO}_3 \text{ will contain } \frac{48 \times 100}{101} = 47.5 \text{ g of O}$$

Hence the percentage of oxygen in $\text{KNO}_3 = 47.5$

Thus the composition of KNO_3 is as follows .

	%
Potassium	38.6
Nitrogen	13.9
Oxygen	47.5
	<hr/>
	100.0
	<hr/>

3 4 Valency and chemical formulae

Valency of atoms

Atoms combine to form molecules. Each atom has a definite capacity to combine with an atom of its own kind or with other atoms. The hydrogen atom is taken as the reference to specify the combining capacity of other elements. Some elements have variable valencies i. e., they exhibit different combining capacities.

Compounds built up from ions are called ionic compounds, e. g., NaCl formed from Na^+ and Cl^- ions. In the crystal of NaCl the oppositely charged Na^+ and Cl^- ions are held together by electrostatic forces.

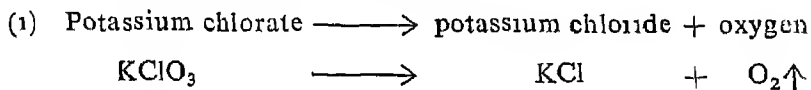
In an ionic compound the total number of positive charges is equal to the total number of negative charges, so that the compound is electrically neutral. On this basis it is possible to write the formulae of ionic compounds as shown in the following table,

Table 1. *Formulae of ionic compounds*

Negative ions	Cl ⁻	I ⁻	S ⁻⁻	SO ₄ ⁻⁻	PO ₄ ⁻⁻⁻
positive ions					
Na ⁺	NaCl	NaI	Na ₂ S	Na ₂ SO ₄	Na ₃ PO ₄
Mg ⁺⁺	MgCl ₂	MgI ₂	MgS	MgSO ₄	Mg ₃ (PO ₄) ₂
Fe ⁺⁺⁺	FeCl ₃	FeI ₃	Fe ₂ S ₃	Fe ₂ (SO ₄) ₃	FePO ₄

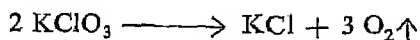
3.5 Chemical equations

To illustrate the steps involved in balancing an equation, you may consider some more reactions like the decomposition of potassium chlorate, the action of hydrochloric acid on sodium carbonate, etc.

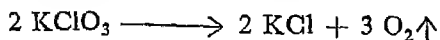


We find that the oxygen atoms are not equal on the two sides of the equation. To equalise the oxygen atoms we multiply the formula KClO₃ by 2 and O₂ by 3.

Thus we get

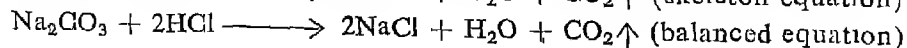
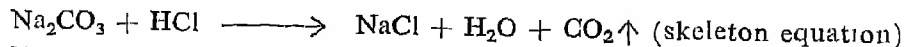


To balance potassium and chlorine multiply KCl by 2



The number of each kind of atoms on both sides of the above equation is equal. Hence, the equation is now balanced.

It is not necessary to follow all the steps suggested in balancing an equation if this could be done by mere inspection, for example, the action of hydrochloric acid on sodium carbonate.



BACKGROUND INFORMATION

3.1 Historical development of symbols

For many centuries a group of people called alchemists attempted (i) to convert base metals like iron and copper into gold; and (ii) to develop medicinal preparations (elixir of life). The alchemists were anxious to keep their knowledge a closely guarded secret. They developed curious symbols to represent the different forms of matter. Some of these symbols were derived from mythology, while some others were associated with heavenly bodies like planets. For example, lance and shield of Mars (God of War) stood for iron, the looking glass of Venus for copper, Moon for silver, Sun for gold, and Saturn for lead.



Figure 3.1 *Alchemical Symbols*

When a large number of substances came to be identified, a more systematic set of symbols became necessary. John Dalton (1766-1844), an English school master, made use of the following symbols.

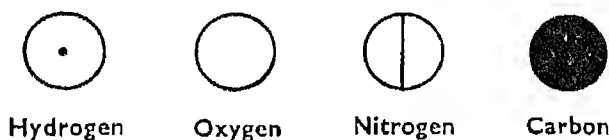


Figure 3.2 *Dalton's symbols for atoms*

To represent compounds, the symbols of the elements present in them were combined, as shown below.



Figure 3.3 *Dalton's symbols for compounds*

The symbols used by Dalton for compounds did not indicate the number of atoms present in them and, therefore, were not appropriate.

Berzelius (a Swedish chemist), in 1814, represented the elements by one or two letters taken from the names of the elements (see page 33, text).

Naming of elements

Some elements are known for many centuries and they are called by their familiar names, e. g., copper, silver, gold and iron.

Elements discovered recently have been named either after the place or country of discovery, or after the name of a scientist. Some examples are :

Polonium (Po) after Poland; Francium (Fr) after France; Californium (Cf) after California, Curium (Cm) after Madam Curie, Einsteinium (Es) after Einstein and Mendelevium (Md) after Mendeleev.

3.3 Empirical and molecular formulae

The formula obtained from the percentage composition of a compound (called empirical formula) gives the ratio of the atoms present in a molecule of the compound. In the case of water, the atomic ratio H : O is 2 : 1. The empirical formula of water is thus H_2O . It could as well be H_4O_2 or $\text{H}_{200}\text{O}_{100}$ etc, but the simplest formula is H_2O . The empirical formula weight of water is 18. The molecular weight of water is also 18. Hence the empirical formula is the same as the molecular formula in this case.

In some cases, however, the empirical formula may not represent the actual number of atoms present in the molecule. This can be illustrated with hydrogen peroxide.

Experimental determination shows that hydrogen peroxide contains 5.88% of hydrogen and 94.12% of oxygen. From this atomic ratio can be calculated.

Element	%Composition	Atomic weight	Atomic ratio
H	5.88	1	$5.88/1 = 5.88$
O	94.12	16	$94.12/16 = 5.88$

The atomic ratio H : O is, therefore, 5.88 : 5.88 or simply 1 : 1. Hence the empirical formula for hydrogen peroxide is HO. The empirical formula weight will then be $1 + 16 = 17$.

The molecular weight of hydrogen peroxide is 34. This value is twice the empirical formula weight. This indicates that the empirical formula is not the same as the molecular formula. The empirical formula (HO) must be multiplied by 2 to get the molecular formula (H_2O_2).

Similarly, from the composition of mercurous chloride (85% Hg, 15% Cl), and its formula weight (473), its formula can be deduced as Hg_2Cl_2 .

3.4 Valency

In the case of some elements, more than one valency is possible. For example, iron forms two chlorides, ferrous chloride (FeCl_2) and ferric chloride (FeCl_3). In FeCl_2 , the valency of iron is 2, whereas in FeCl_3 , its valency is 3. Another example of an element which exhibits variable valency is phosphorus. It forms phosphorus trichloride (PCl_3) and phosphorus pentachloride (PCl_5) with valencies of phosphorus 3 and 5 in the two compounds, respectively.

The maximum valency shown by any element is 8. For example, the valency of osmium in osmium tetroxide, OsO_4 , is 8.

Modern views on valency

In the text, valency has been defined as the combining capacity in terms of the number of atoms of hydrogen with which one atom of an element combines. But a knowledge of electronic configuration of atoms provides a better understanding of valency.

An atom consists of a positively charged nucleus, consisting of positively charged protons and neutral particles called neutrons. The nucleus is surrounded by negatively charged electrons. The number of electrons is equal to the number of protons in the nucleus, so that the atom as a whole is electrically neutral.

The electrons in an atom are arranged in various shells. The maximum number of electrons that can go into each shell is given by $2n^2$, where n is the shell number.

Shell No.	Max. no of electrons
1	$2 \times 1^2 = 2$
2	$2 \times 2^2 = 8$
3	$2 \times 3^2 = 18$
4	$2 \times 4^2 = 32$

The number of protons in an element is the atomic number. Naturally, in the neutral atom, the number of electrons is also equal to the atomic number.

In the periodic table the elements are arranged in the order of increasing atomic numbers. For example, the atomic number of sodium is eleven because it contains eleven protons in its nucleus. It is, therefore, the eleventh element in the periodic table. In the neutral sodium atom, the nucleus is surrounded by eleven electrons. Similarly chlorine whose atomic number is 17 will have 17 electrons when the atom is neutral. These electrons are arranged in three shells as 2, 8, 7. The arrangement of electrons in various shells in the atoms of some elements is shown below :

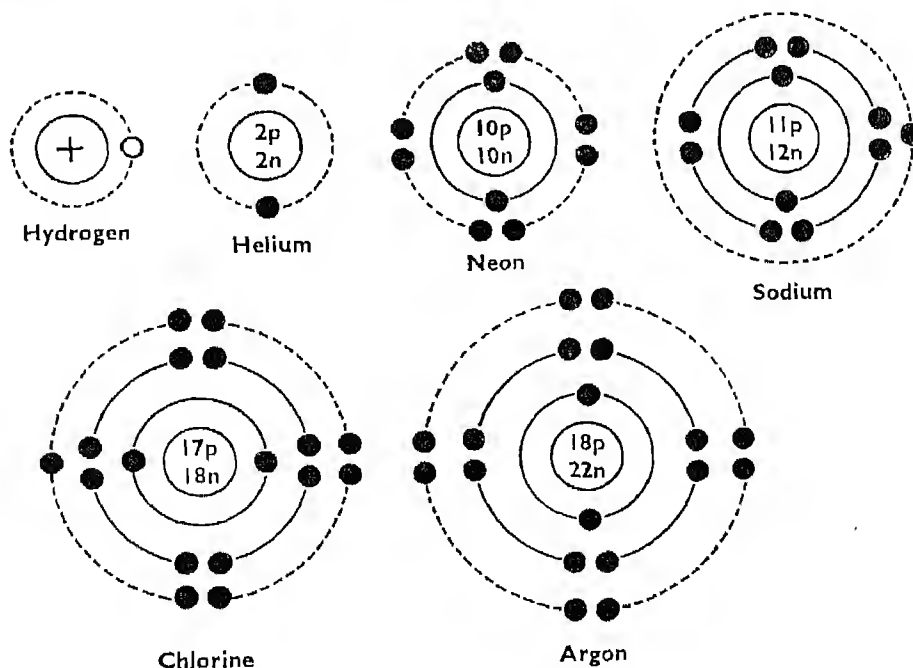


Figure 3.4 Electronic arrangement in some atoms

3.5 (i) Ionic valency

Generally atoms of different elements have a tendency to lose or gain electrons. This tendency varies from element to element. Metallic elements like sodium, potassium or magnesium have a tendency to lose one or more electrons forming thereby the positive ions Na^+ , K^+ , or Mg^{++} . On the other hand some elements like chlorine or oxygen have a tendency to gain one or more electrons forming thereby negative ions Cl^- or O^{--} . By such a loss or gain of electrons the atoms of these elements acquire the electronic structure of the nearest inert gas.

For example, the sodium ion, Na^+ , has 11 protons in the nucleus, but only 10 electrons outside the nucleus. This accounts for the unit positive charge of the sodium ion. In this form the sodium ion, Na^+ , resembles the neon atom as both have 10 electrons.

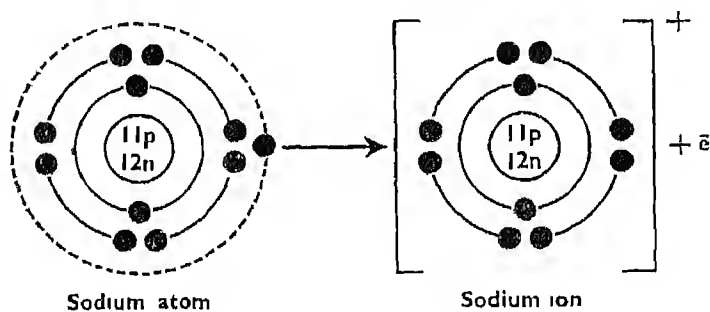


Figure 3.5 *Formation of sodium ion*

Similarly the chlorine atom (Cl) has one electron less than an atom of argon. So chlorine will have a tendency to take up an electron and form chloride ion (Cl^-), which will have the electronic structure of argon.

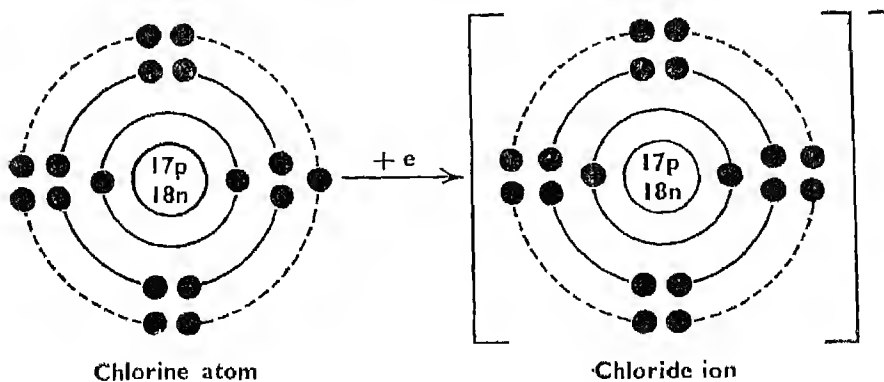


Figure 3.6 *Formation of chloride ion*

When sodium reacts with chlorine, sodium chloride is formed. In this reaction sodium atom loses an electron and forms a sodium ion, Na^+ . The electron given by sodium atom is taken up by chlorine atom forming a chloride ion, Cl^- . The Na^+ and Cl^- ions thus formed are held together by electrostatic forces in the sodium chloride crystal. A crystal of sodium chloride contains equal number of Na^+ and Cl^- ions. These ions are closely packed in the crystal in an orderly manner as shown below.

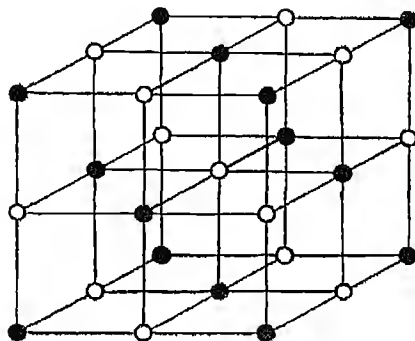


Figure 3.7 Packing of ions in sodium chloride crystal

The formation of an ion with more than one positive or negative charge can be explained thus :

Magnesium has the electronic structure 2, 8, 2. By giving up the two electrons of the outer-most shell, Mg becomes Mg^{++} ion with the electronic structure of neon (2, 8).

Oxygen atom with the electronic structure 2, 6 can take up 2 electrons to form the oxide ion, O^{--} , with the electronic structure of neon (2, 8). Magnesium ions (Mg^{++}) and oxide ions (O^{--}) are held together by electrostatic force in magnesium oxide, MgO .

Thus many atoms are capable of losing or gaining electrons to form positive and negative ions which have the stable electronic structures of the inert gas atoms. The positive and negative ions thus formed are held together by electrostatic forces in ionic compounds.

(ii) Covalency

Hydrogen and chlorine combine to form hydrogen chloride. Hydrogen chloride is a gas with hardly any ions in it. What then are the forces which hold hydrogen and chlorine together?

Let us consider the electronic structure of hydrogen and chlorine

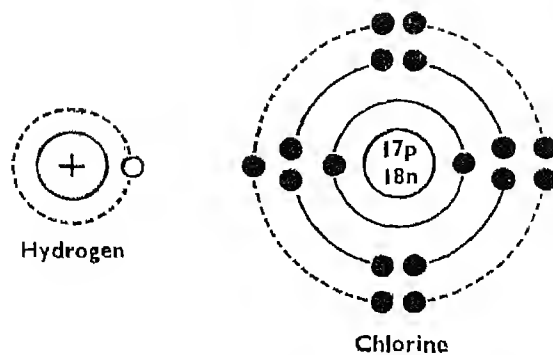


Figure 3.8 *Structure of hydrogen and chlorine atoms*

Hydrogen atom has one electron less than the helium atom. Chlorine atom has also one electron less than the argon atom. However, both hydrogen and chlorine can acquire an inert gas configuration by mutual sharing of electrons.

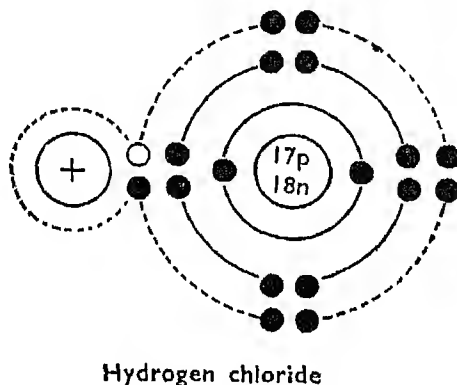


Figure 3.9 *Formation of hydrogen chloride*

This shared pair of electrons holds the two atoms together and constitutes a bond. This bond is called covalent bond.

In water, H_2O , there are two covalent bonds between the oxygen and hydrogen atoms as shown below

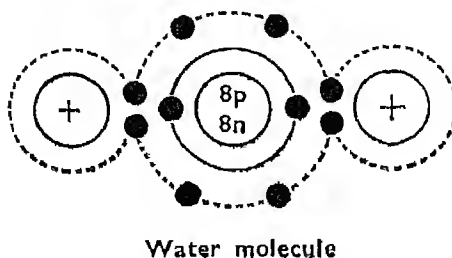


Figure 3.10 *Formation of water molecule*

The covalent bond is usually represented by a dash (—) between the combining atoms

What is meant by the symbol C?

What do the following formulae indicate?

2H_2 , CO_2

What is valency?

Which element is considered as a reference for comparison of valencies and why?

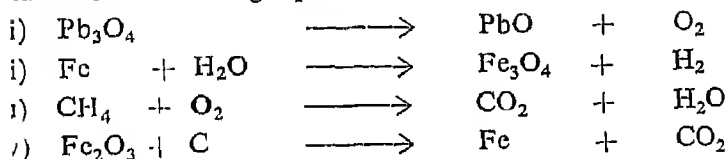
Give examples of monovalent, divalent and trivalent elements.

Elements X and Y form compounds with carbon, the molecules of which can be represented by the formulae CX_4 and CY_2 . Predict the formula of a compound of X and Y.

Elements X, Y, Z and H form the following compounds: HX , YH , YZ . Deduce the valencies of X, Y and Z (Valency of H is one).

A compound of hydrogen and carbon contains 7.7% hydrogen. What is the simple formula? If the molecular weight is 78, what is the molecular formula?

Balance the following equations.



How many atoms are present in

- (i) 28 g of nitrogen
- (ii) 8 g of sulphur
- (iii) 64 g of oxygen?

How much would the following weigh?

- (i) 1.5×10^{23} atoms of calcium
- (ii) 12×10^{23} atoms of sodium
- (iii) 3×10^{23} atoms of magnesium.

Solutions

INSTRUCTION MATERIAL

Introduction

Solutions play an important role in life processes as well as in chemical reactions. A study of solutions is, therefore, an important aspect of chemistry.

In this unit an attempt is made to make the pupils understand the various types of solutions and the factors which affect the rate of dissolution of a solute. Electrical conductivity of aqueous solutions is introduced to familiarise the pupils with the concept of ions. The process of crystallisation is also dealt with.

Outline

1. What is a solution ?(4.1).
2. Process of dissolution(4.2).
3. Saturated and unsaturated solutions (4.3).
4. Electrical conductivity of aqueous solutions (4.4).
5. Types of solutions (4.5).

Concepts

1. A solution is a homogeneous mixture of two or more substances.
2. The common type of solution involves a solid (solute) dissolved in a liquid (solvent). Water is a widely used solvent.
3. There is a limit to the amount of solute that can be dissolved in a given quantity of a solvent at a given temperature.
4. Some substances (electrolytes), when dissolved in water, dissociate into positive and negative ions.

SCHEDULE

Lesson	Experiment	Postlab	Demonstration	Text covered
1	10 and 11	Discussion	I	4.1
2	12	Discussion	II	4.1 and 4.2
3	13 and 14	—	III	4.3
4	15	Discussion	IV	4.4 and 4.5
5	16	—	V and VI	4.5
6	—	—		Review and quiz

Development

4.1 What is a solution ?

A study of solutions begins with experiment 10. This experiment shows that a solution is a homogeneous mixture of two substances. Explain to the pupils that in a homogeneous mixture the composition is uniform.

Experiment 10 Preparation of a solution (solid in liquid)

Time required : 15 minutes

Questions and answers

1. What is the colour of potassium permanganate crystals ?
Deep purple.
2. What happened immediately after the crystals were put in water ?
As the crystals moved down the water layer they left behind a purple trail.
3. Was the colour of the solution uniform after five minutes ?
No.
4. Was the colour of the solution uniform after stirring ?
Yes.

After knowing how to prepare solutions, the pupils may be led to know that energy changes usually accompany the process of dissolution. Experiment 11 illustrates these changes.

Experiment 11 Heat changes accompanying dissolution of substances in water

Time required : 20 minutes

It may be difficult to provide each pupil with a thermometer. So the heat changes are noted in a qualitative manner by touching the beaker in which dissolution is occurring. You may, however, measure the heat changes using a thermometer.

Postlab discussion

Explain the meaning of the terms 'endothermic' and 'exothermic'. When solid ammonium chloride goes into solution heat is absorbed from the solvent and the temperature of the solution decreases (draw the attention of the pupils to the formation of dew on the outer surface of the beaker). In a similar manner, when sodium hydroxide goes into solution, heat is liberated. This heat is taken up by the solvent and the temperature of the solution increases. When sodium chloride is dissolved in water there is no appreciable change in temperature.

Mention may be made here of the preparation of dilute sulphuric acid. The concentrated acid should be added slowly, with stirring, to a large amount of water. This reaction is highly exothermic and should be carried out with care. (*Caution: Water should never be added to concentrated sulphuric acid. In the process of dissolution of the acid in water, the heat liberated is so large that the small amount of water added is converted to steam and the acid spurts*).

As an example of an exothermic reaction, you may cite the preparation of slaked lime (used in white washing) by adding quick lime to water. In this process much heat is liberated and sets the water boiling.

It should be emphasised that the components of a solution (solute and solvent) can be separated. This justifies the statement made in unit 1 that a solution is also a kind of mixture. Demonstration I serves this purpose.

Demonstration I Recovery of sodium chloride from its solution

Time required : 30 minutes

Materials required

Evaporating dish (100 ml)

Sodium chloride 10 g

Funnel

Filter paper

Sand bath

Tripod

Funnel stand

Spirit lamp

Glass rod

Measuring cylinder (50 ml)

Box of matches

Procedure

Prepare a solution of sodium chloride by dissolving 10 g of the salt in 50 ml of water (filter, if necessary).

Evaporate the solution in an evaporating dish using a sand bath. Towards the end, use a small flame to avoid spurting. When all the water has evaporated, cool the dish and collect the residue on a paper.

During evaporation the solvent escapes into the atmosphere, leaving the solute behind. The solvent can be recovered by cooling its vapour in a water cooled condenser. Show this by demonstration II.

Demonstration II Recovery of water and sodium chloride from a solution of the salt

Time required : 40 minutes

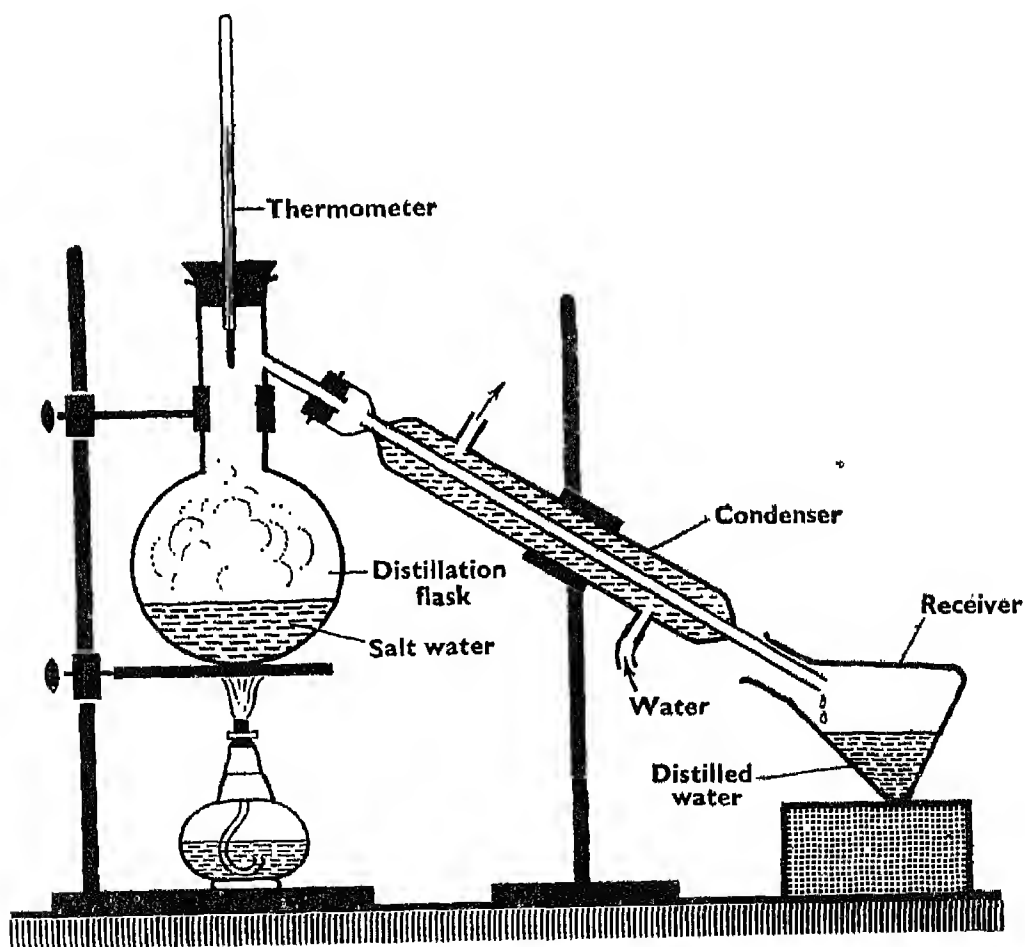


Figure 4.1 Distillation apparatus

Materials required

Distillation flask (100 ml)	Sodium chloride 10 g
Liebig's condenser	
Thermometer	
Wire gauze	
Iron stands with adjustable clamps	2
Sand bath	
Spirit lamp	
Rubber tubing	
Corks	2
Cork borer set	
Wooden block	
Measuring cylinder (50 ml)	
Conical flask	
Box of matches	

Procedure

Dissolve 10 g of sodium chloride in 50 ml of water. Take the solution in a distillation flask fitted with a one holed cork carrying a thermometer. Connect the side tube of the flask to a condenser through which cold water is circulated as shown in figure 4.1 (before you start heating the solution, explain the function of each part used). Heat the flask and collect the distillate in a receiver.

When the volume of the solution is reduced to about 5 to 10 ml, transfer it to an evaporating dish and evaporate it to dryness. Cool the dish and transfer the residue to a paper.

4.2 Process of dissolution

After learning the process of dissolution and the separation of the components of a solution, the pupils may be asked to perform experiment 12.

Experiment 12 Factors affecting the rate of dissolution

Time required: 30 minutes

Postlab discussion

Elicit from the pupils the fact that three factors — powdering, stirring and heating, increase the rate of dissolution. The influence of these factors may be illustrated with some examples chosen from experiences in daily life.

Questions and answers

1. In which beaker did the candy dissolve first ?
In beaker B.

2. What is the effect of stirring on the rate of dissolution of powdered candy?

To increase the rate of dissolution.

4.3 Solubility of salt in water

There is a limit to the extent to which a salt dissolves in water at a given temperature. When this limit is reached, a saturated solution is formed. This may be shown by demonstration III.

Demonstration III Preparation of a saturated solution of sodium chloride

Time required : 20 minutes

Materials required

Beaker (100 ml)

Sodium chloride (powdered) 25 g

Glass rod

Measuring cylinder (50 ml)

Procedure

Prepare five packets of sodium chloride each containing 5 g. Take 50 ml of water in a 100 ml beaker. Add 5 g of powdered sodium chloride and stir. Does the salt dissolve completely?

To find out if this solution is saturated or not, add another 5 g of sodium chloride and stir. Does the second portion of the salt go into solution?

Repeat the above procedure till some salt remains undissolved, i. e., a saturated solution is obtained. Emphasise that the solubility varies from solute to solute in a given solvent at a particular temperature.

After this demonstration the pupils may be asked to do experiment 13. For use in the laboratory pure substances are required. Many of the commercial samples of substances are not pure. Experiment 13 describes a simple and convenient method of obtaining a pure sample from a commercial one.

Experiment 13 Purification of a commercial sample of copper sulphate

Time required . 30 minutes

The crystals obtained are small in size. Large crystals may be obtained by allowing the solution to cool slowly. By suspending a small crystal in a saturated solution of the substance a large crystal may be made to grow. This may be assigned to the pupils as an activity.

Experiment 14 Preparation of solutions of different concentrations

Time required : 10 minutes

Mention may be made of the preparation of dilute acids (hydrochloric and sulphuric acids) by mixing concentrated acids with water.

4.4 Electrical conductivity of aqueous solutions

Aqueous solutions of some substances conduct electricity, while those of some others do not. The former are known as electrolytes and the latter non-electrolytes. Demonstration IV helps in knowing which among the substances taken are electrolytes and which, non-electrolytes.

Demonstration IV Electrical conductivity of aqueous solutions of a few substances

Time required : 25 minutes

Materials required

Beaker (100 ml)	Copper sulphate	1 g
Glass rod	Cane sugar	1 g
Copper strips (10 x 1 x 0.2 cm) 2	Sodium chloride	1 g
Card board (7 cm square)	Distilled water	500 ml
Torch bulb (2.5 V)	Ethyl alcohol	50 ml
Battery (6 V)	Dilute sulphuric acid	50 ml
Plug-key	Dilute sodium hydroxide solution	50 ml

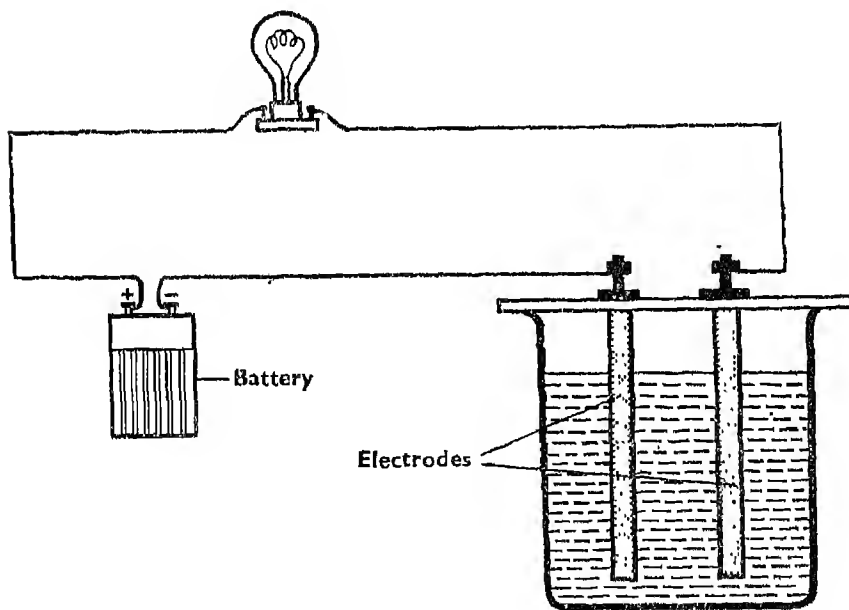


Figure 4.2 Conductivity of solutions

Procedure

Connect a battery, a bulb, a plug-key and two copper strips as shown in figure 4.2. Dip the copper strips in distilled water in a 100 ml beaker. Close the plug-key and see whether the bulb glows or not. Replace the water in the beaker

by a solution of 1 g of copper sulphate in 50 ml of water. Close the plug-key. See if the bulb glows now. Repeat with solutions of sodium chloride, sulphuric acid, sodium hydroxide, alcohol and sugar in water.

Discussion

The bulb does not glow when the copper strips are dipped in distilled water, sugar and alcohol solutions. But it glows when they are dipped in solutions of copper sulphate, sodium chloride, sulphuric acid and sodium hydroxide. Explain why some solutions conduct electricity while others do not.

4.5 Types of solutions

Water is a good solvent for many substances. There are however some substances which do not dissolve in water, for example, paraffin wax. This substance dissolves in kerosene. In experiment 15, the pupils will test the solubility of some substances in a few solvents.

Experiment 15 The solubility of solids in various liquids

Time required : 25 minutes

Postlab discussion

On the basis of this experiment, a broad classification into aqueous and nonaqueous solutions may be made. Solutions may also be obtained by dissolving one liquid in another. Rose water is an example of the solution of a liquid in liquid. Such liquid in liquid solutions are dealt with in experiment 16.

Experiment 16 Miscibility of liquids

Time required : 20 minutes

This experiment shows that some liquids are miscible with one another, for example, water and spirit, while others are not, for example, water and kerosene.

Just as solids and liquids dissolve in liquids, gases also dissolve in liquids, for instance, air dissolves in water to a slight extent. Show its presence in water by means of demonstration V.

Demonstration V Water contains dissolved air

Time required : 15 minutes

Materials required

- Beaker (250 ml)
- Funnel (with short stem)
- Test tube
- Wire gauze
- Tripod
- Spirit lamp
- Box of matches

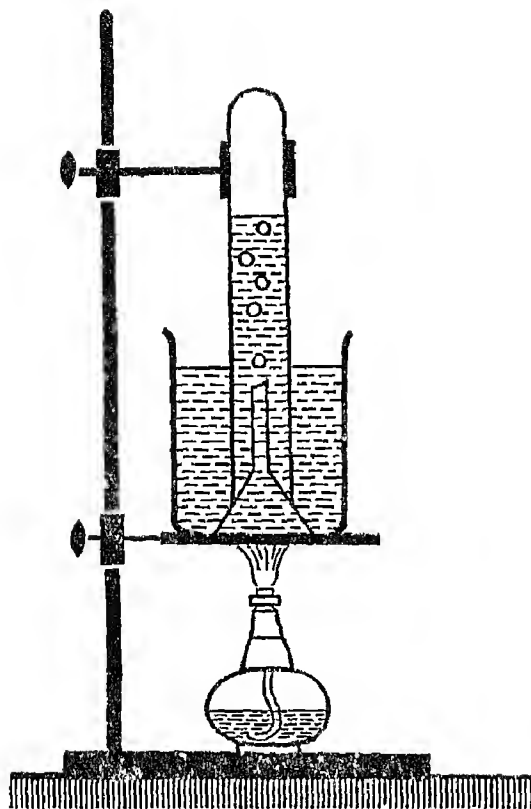


Figure 4.3 *Solubility of air in water*

Procedure

Take 200 ml of water in a 250 ml beaker. Place a short stemmed funnel in the inverted position as shown in figure 4.3. See that the stem of the funnel is completely immersed in water. Invert a test tube filled with water over the stem of the funnel. Heat the water in the beaker to about 60°C . Notice the appearance of air bubbles which collect in the test tube.

Discussion

The gas collected in the test tube may be shown to be air by introducing a burning splinter which continues to burn. Draw the attention of the pupils to the formation of small air bubbles on the sides of a bucket containing water, when the bucket is exposed to sunlight for a long time.

The solubility of air in water, is not high. But there are some gases which are highly soluble in water, for example, hydrogen chloride and ammonia. The high solubility of hydrogen chloride in water may be shown by demonstration VI.

Demonstration VI Solubility of hydrogen chloride in water

Time required . 15 minutes

Materials required

Beaker (250 ml)

Blue litmus solution 300 ml

Round bottomed flask (250 ml)

Hydrogen chloride generator

Glass tube with tapering end

Dropper

Glass tubes

Rubber cork with two holes

Iron stand with ring

Test tube

Procedure

Prepare hydrogen chloride gas by the action of concentrated sulphuric acid on sodium chloride. Collect the gas generated in a test tube. Close the test tube with your thumb and invert it in a beaker of water. Remove the thumb and show

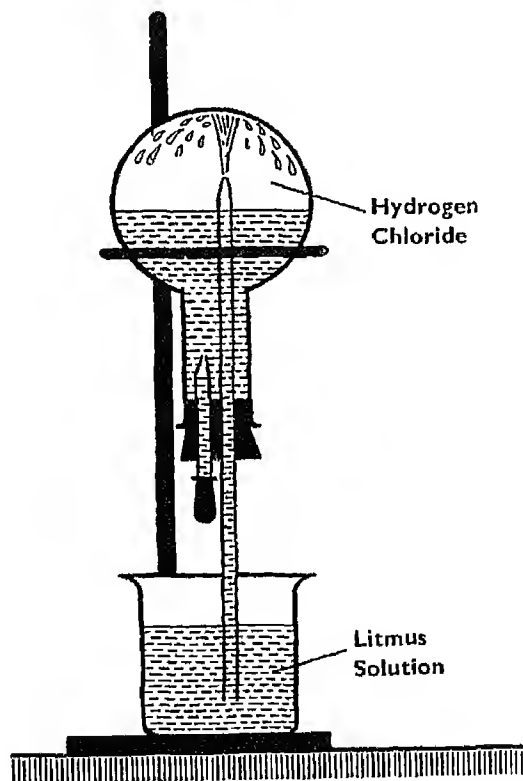


Figure 4.4 Solubility of hydrogen chloride in water

that the water rushes in and fills the test tube, indicating that hydrogen chloride is highly soluble in water. The same property may be demonstrated more spectacularly as follows :

Collect hydrogen chloride gas in a dry, 100 ml round bottomed flask. Fit the flask with a two holed rubber cork. Insert a dropper containing a small quantity of water in one hole and a narrow glass tube tapering at one end in the other. Invert the flask such that the glass tube dips in blue litmus solution contained in a beaker, as shown in figure 4.4.

Squeeze the dropper gently so that a few drops of water enter the flask. Immediately the blue litmus solution rushes into the flask in the form of a fountain. The blue litmus solution turns red.

Explain why the blue litmus solution rushes into the flask on squeezing the dropper. Emphasise that the change in colour of the litmus solution has nothing to do with the solubility of hydrogen chloride in water. The change is due to the characteristic property of acids. They will learn about this in unit 8.

Draw the attention of the pupils to the high solubility of carbon dioxide and ammonia in water.

BACKGROUND INFORMATION

4.1 What is a solution?

Heat changes during dissolution

When solid ammonium chloride goes into solution, heat is absorbed (endothermic dissolution) and when solid sodium hydroxide dissolves in water, heat is liberated (exothermic dissolution).

The magnitude of the heat changes vary with the type of substance, its quantity and the amount of water used. In this process the energy required to break the ionic bonds in the crystal (lattice energy) is supplied by the system (water+ion) and the temperature is lowered. However, these ions and water molecules further react and form hydrated species. This hydration is accompanied by the liberation of energy (hydration energy) and the system liberates heat.

The ability of water to dissolve ionic crystals is best explained on the basis of the polar nature of water (page 36).

4.2 Process of dissolution

The phenomenon of dissolution of a solute in a solvent can be explained on the basis of one or more of the following processes.

- i. Chemical reaction with the solvent to form a new substance

Calcium phosphate reacts with dilute hydrochloric acid to form calcium chloride and phosphoric acid



- ii. Chemical reaction with a solvent to form a solvated substance

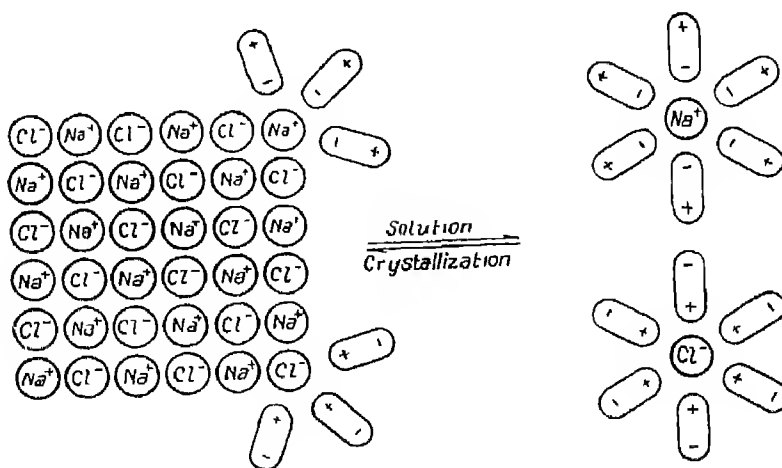


Figure 4.5 Dissolution of an ionic compound

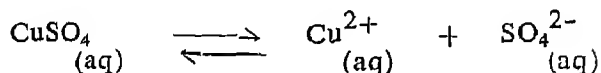
Such a solvation process is understood on the basis of the polar nature of a solvent, like for instance, water (page 33). A crystal of sodium chloride contains sodium ions Na^+ and chloride ions, Cl^- , held together by electrostatic forces. If now a crystal of sodium chloride is placed in water, the positive end of the water molecule is attracted to the negative chloride ion, which is surrounded by a cluster of water molecules. Similarly the negative end of the water molecule is attracted to the positive sodium ion, which also is clustered by water molecules. This is known as "hydration" of the ions. The hydration process supplies the necessary energy to overcome the electrostatic attractions between sodium and chloride ions

111. Dissolution due to dispersion

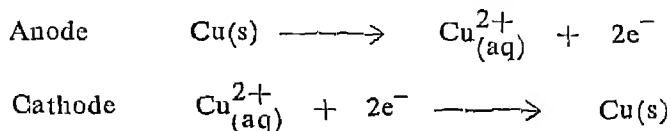
In such cases the nature and magnitude of the forces between solute molecules on the one hand and solvent molecules on the other are more or less the same. As such the solute merely drifts into the solvent to form a solution. However, it should be noted that the process of dissolution cannot be explained by any single comprehensive mechanism; usually more than one mechanism is involved.

4.4 Electrical conductivity of aqueous solutions

The electrical conductivity of aqueous copper sulphate is due to its dissociation into positive ions (cations) and negative ions (anions) as shown below:

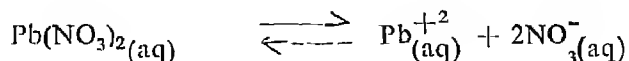
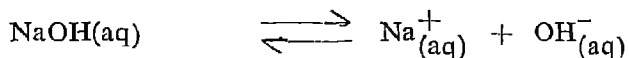
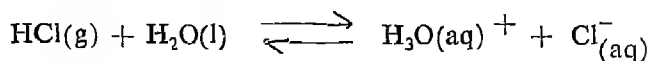


When an electric current passes through the copper sulphate solution, electrolysis takes place. The copper strips connected to the positive and negative terminals of the battery act as the anode and the cathode, respectively. A small quantity of copper dissolves from the anode and gets deposited on the cathode (in demonstration IV a red deposit is seen on the cathode). During the electrolysis the following changes occur at the electrodes.



The role of water during the dissolution of salts like sodium chloride and copper sulphate is only to bring about the separation of the ions which already exist in the solid. In sodium chloride crystals, the Na^+ and Cl^- ions are held together by electrostatic forces and so are not free to move about. When the solid is dissolved in water the crystal structure is destroyed and the ions move freely in the solution. The free movement of charged ions enables the current to pass through the solution. (When sodium chloride is melted the crystal structure is destroyed and hence the molten salt is also able to conduct electricity).

All acids, bases and salts undergo dissociation to a smaller or larger extent when they dissolve in water. Some examples are given below



Quiz

1. Fill up the blanks by choosing the correct answer.
 - a. Solution is obtained by dissolving a solid in a
(i) solution (ii) solvent (iii) solute
 - b. Soda water is a solution of in water.
(i) air (ii) carbon dioxide (iii) oxygen
 - c. Sand is in water.
(i) soluble (ii) insoluble
 - d. Stirring the rate of dissolution of sugar in water.
(i) has no effect on (ii) decreases (iii) increases
 - e. To get a dilute solution of copper sulphate, we should
the concentrated solution.
(i) boil (ii) add water to (iii) add copper sulphate to
 - f. Common salt is in kerosene.
(i) soluble (ii) insoluble
 - g. Common salt is obtained by sea water.
(i) freezing (ii) filtering (iii) evaporating
2. Select suitable solvents for dissolving the following solutes.
 - a. Iodine
 - b. Ammonia
 - c. Sugar
 - d. Sulphur

(Solvents : water, alcohol, carbon disulphide, kerosene)
3. Complete the following statements with one or more suitable words.
 - a. A solution is obtained by dissolving a in a solvent.
 - b. An aqueous solution is prepared by using as the solvent.

- c. Stirring ... the rate of dissolution of salts in water
 - d. The process of formation of crystals is called ...
 - e. The shape of sodium chloride crystal is
 - f. Copper sulphate is called an since its aqueous solution conducts electricity.
4. State whether the following statements are True or False (Write 'T' for true, 'F' for false).
- a. A solution is a homogeneous system. ()
 - b. Powdered sugar dissolves faster than sugar candy. ()
 - c. When ammonium chloride is dissolved in water, heat is evolved. ()
 - d. A solution of cane sugar in water conducts electricity. ()
 - e. An unsaturated solution dissolves more of the solute. ()
 - f. Iodine dissolves readily in water but not in alcohol. ()
 - g. Soda water is a solution of oxygen in water. ()

CHEMICALS AND EQUIPMENT FOR THE LABORATORY EXPERIMENTS

The following lists serve as a guide for the purchase of chemicals and equipment required for one academic year. Chemicals are consumed and stocks are to be replenished. Equipment consists of items which can be reused, such as, glass apparatus and hardware. It may not be possible to supply some items of the equipment individually to each pupil. In such cases only a few may be stocked and used by the pupils by turns.

Chemicals

The following list indicates the amounts required for a class of 40 pupils, each pupil working individually as far as possible.

Substance	Quantity	Used in experiment
Ammonium chloride	60 g	3,11
Candies (about 1/2 cm cube)	240 pieces	12
Calcium chloride (anhydrous)	200 g	4
Chalk powder	40 g	15
Copper sulphate	240 g	3,13,14,15
Iodine	40 g	15
Kerosene	1200 ml	15,16
Lead nitrate	20 g	3
Magnesium ribbon (4 cm)	40 pieces	3
Naphthalene	40 g	15
Nichrome wire (4 cm)	40 pieces	3
Paraffin wax	40 g	15
Paraffin oil	40 ml	4
Potassium permanganate	4 g	10
Sand	120 g	4
Sodium chloride	160 g	5,11,15
Sodium hydroxide	40 g	11
Spirit (methylated)	8 liters	2,3,5,8,12,13,15,16
Vegetable oil	400 ml	16
Asbestos mat (19 x 15 cm)	40	5,13

Equipment	Quantity	Used in experiment
Beakers (100 ml)	80	1,5,10,13
Beakers (250 ml)	80	12
Box of matches	40	2,3,4,5,8,12,13
Corks (to fit test tubes)	240	1,15,16
Cotton	40 g	5
Evaporating dishes (100 ml)	40	5,13
Filter stands	40	5,13
Filter paper circles	80	5,13
Funnels	40	2,5
Glass rods (20 cm)	80	1,5,10,11,12,13,14
Graduated test tubes	40	16
Hand lens	40	1
Iron nails (1cm)	360	4
Iron stands with clamp	40	8
Measuring cylinders (100 ml)	40	12
Mortar and Pestles	40	12
Pair of tongs	40	3,5,13
Plastic combs	40	9
Rubber balloons	40	6
Rubber tubings (25 cm)	40 pieces	7
Rubber stoppers carrying a delivery tube	40	8
Spirit lamps	40	2,3,4,8,12,13
Syringes (10 ml)	40	7
Terylene or silk cloth (25 cm sq.)	40 pieces	9
Test tubes (hard glass)	120	3,8
Test tubes (ordinay)	240	1,4,11,14,15,16
Test tube brushes	40	15
Test tube holders	40	3,4
Tripods	40	5,32,13
Troughs	40	8
'U' tubes	40	7
Wash bottles (100 ml)	40	5
Wire gauzes	40	5,12,13

EQUIPMENT REQUIRED FOR DEMONSTRATION EXPERIMENTS

The following is the list of chemicals and apparatus needed for showing demonstration experiments of units 1-4, *for one section only*. The chemicals which are consumables have to be replenished if the experiments are to be demonstrated to another section. The apparatus, however, could be used any number of times.

APPARATUS

Name	Quantity
Balance	1
Balloons (different sizes)	12
Bamboo stick 30cm \times 0.5cm (diameter)	1
Battery 6 V.	1
Beaker 150 ml.	1
Bell jar with stopper	1
Boiling tube	1
Box of matches	1
Bulb 2.5 V.	1
Burette 50 ml.	1
Card board 7 cm \times 7 cm	1
Capillary tube 50 cm \times 0.1 cm (bore)	1
Condenser (Liebig's)	1
Copper strips 10 cm \times 1 cm \times 0.2 cm	2
Cork borer set	1
Corks (assorted sizes)	12
Cotton thread	50 cm
Cup voltmeter	1
Cycle valve tube	2 cm
Deflagrating spoon	1
Delivery tube	1
Distillation flask 100 ml	1
Droppers	2
Ebonite rod	1
Evaporating dish 100 ml	1
Filter papers	2
Flask (Flat Bottom) 150 ml	1
(Round Bottom) 250 ml	1
Funnel	1
Funnel stand	1
Gas collecting jars	2
Glass bottle	1
Glass marking pencil	1
Glass plates 10 cm \times 10 cm	2

Name		Quantity
Glass tube 30 cm \times 2.5 cm (diameter)	.	1
Glass tube 10 cm bent at right angles	...	1
Glass tube 50 cm \times 0.6 cm (diameter)	..	1
Glass tube with tapering end 25 cm	..	1
Glass rod 1 cm \times 0.7 - 0.8 cm	..	1
Glass rods	..	2
Hand lens	...	1
Ignition tube	...	1
Inflator	..	1
Iron Stand	...	1
Clamps	...	2
Boss heads	..	2
Iron ring	..	1
Knife	...	1
Marbles	..	2
Magnet	...	1
Measuring cylinder 100 ml	...	1
Paper weight	..	1
Plug-key	...	1
Rubber corks (2 holed) to fit the boiling-tube	..	2
Rubber tubing	...	3½ metres
Sand bath	...	1
Scale (1 meter)	...	1
Spirit lamp	...	1
Syringe 10 ml	...	1
Test tubes	...	6
Test tube (Hard glass)	..	1
Terylene cloth 15 cm \times 15 cm	...	1
Tripod	...	1
Trough	...	1
Thermometer 0°-110°C	..	1
Watch glass 5 cm	...	3
10 cm	...	1
Wire gauze	..	1
CHEMICALS		
Acid acetic	...	2 ml
Acid hydrochloric	...	15 ml
Acid nitric	...	8 ml
Acid sulphuric	..	20 ml
Alcohol	...	50 ml
Carbon disulphide	...	20 ml
Carbon tetrachloride	..	10 ml
Copper turnings	..	3 g

Name			Quantity
Copper sulphate	..	.	1 g
Iron filings	13 g
Iron wire			4 cm
Iron wire (stout)	20 cm
Lead nitrate	..		1 g
Liquor ammonia	4 ml
Litmus	.	..	2 g
Manganese dioxide	1 g
Mercury		..	2 ml
Phosphorous (yellow)	.	.	0.5 g
Potassium iodide	1.0 g
Sulphur		...	10 g
Splinters	..	.	6
Sodium chloride	.	.	50 g
Sodium metal	1 g
Sodium hydroxide		...	1 g
Sugar	1 g
Sodium thiosulphate	..	.	2 g
Spirit	2 litres

International Atomic Weights

Element	Symbol	Atomic No.	Atomic Wt.	Element	Symbol	Atomic No.	Atomic Wt.
Actinium	Ac	89	227	Einsteinium	Es	99	[254]
Aluminium	Al	13	26.98	Erbium	Er	68	167.27
Americium	Am	95	[243]*	Europium	Eu	63	152.0
Antimony	Sb	51	121.76	Fermium	Fm	100	[253]
Argon	Ar	18	39.994	Flourine	F	9	19.00
Arsenic	As	33	74.91	Francium	Fr	87	[223]
Astatine	At	85	[210]	Gadolinium	Gd	64	157.26
Barium	Ba	56	137.36	Gallium	Ga	31	69.72
Berkelium	Bk	97	[249]	Germanium	Ge	32	72.60
Beryllium	Be	4	9.013	Gold	Au	79	197.0
Bismuth	Bi	83	209.00	Hafnium	Hf	72	178.50
Boron	B	5	10.82	Helium	He	2	4.003
Bromine	Br	35	79.916	Holmium	Ho	67	164.94
Cadmium	Cd	48	112.41	Hydrogen	H	1	1.008
Calcium	Ca	20	40.08	Indium	In	49	114.82
Californium	Cf	98	[251]	Iodine	I	53	126.91
Carbon	C	6	12.011	Iridium	Ir	77	192.2
Cerium	Ce	58	140.13	Iron	Fe	26	55.85
Cesium	Cs	55	132.91	Krypton	Kr	36	83.80
Chlorine	Cl	17	35.457	Lanthanum	La	57	138.92
Chromium	Cr	24	52.01	Lead	Pb	82	207.21
Cobalt	Co	27	58.94	Lithium	Li	3	6.940
Copper	Cu	29	63.54	Lutetium	Lu	71	174.99
Curium	Cm	96	[247]	Magnesium	Mg	12	24.32
Dysprosium	Dy	66	162.51	Manganese	Mn	25	54.94

*Values in brackets are mass numbers of longest-lived or best-known isotopes

Element	Symbol	Atomic No	Atomic Wt	Element	Symbol	Atomic No	Atomic Wt
Mendelevium	Md	101	[256]	Ruthenium	Ru	44	101.1
Mercury	Hg	80	200.61	Samarium	Sm	62	150.35
Molybdenum	Mo	42	95.95	Scandium	Sc	21	44.96
Neodymium	Nd	60	144.27	Selenium	Se	34	78.96
Neon	Ne	10	20.183	Silicon	Si	14	28.09
Neptunium	Np	93	[237]	Silver	Ag	47	107.880
Nickel	Ni	28	58.71	Sodium	Na	11	22.991
Niobium	Nb	41	92.91	Strontium	Sr	38	87.63
(Columbium)				Sulphur	S	16	32.066
Nitrogen	N	7	14.008	Tantalum	Ta	73	180.95
Osmium	Os	76	190.2	Technetium	Tc	43	[99]
Oxygen	O	8	16	Tellurium	Te	52	127.61
Palladium	Pd	46	106.4	Terbium	Tb	65	158.93
Phosphorus	P	15	30.975	Thallium	Tl	81	204.39
Platinum	Pt	78	195.09	Thorium	Th	90	232.05
Plutonium	Pu	94	[242]	Thulium	Tm	69	168.94
Polonium	Po	84	210	Tin	Sn	50	118.70
Potassium	K	19	39.100	Titanium	Ti	22	47.90
Praseodymium	Pr	59	140.92	Tungsten	W	74	183.86
Promethium	Pm	61	[145]	Uranium	U	92	238.07
Protactinium	Pa	91	231	Vanadium	V	23	50.95
Radium	Ra	88	226.05	Xenon	Xe	54	131.30
Radon	Rn	86	222	Ytterbium	Yb	70	173.04
Rhenium	Re	75	186.22	Yttrium	Y	39	88.92
Rhodium	Rh	45	102.91	Zinc	Zn	30	65.38
Rubidium	Rb	37	85.48	Zirconium	Zr	40	91.22

Dear Reader

The present experimental editions have been brought out as a result of the Education Commission's recommendation that Chemistry should be taught as a separate discipline in our schools from class 6 onwards. This has become necessary in view of the great strides that chemistry has made in recent years and the consequent need for the pupils in our schools to be exposed to the modern ideas and trends in the subject. This modest effort, it is hoped, will fulfil that long felt need. It is aimed at presenting the subject in a simple language so as to be comprehensible to pupils in the age group 10 plus to 12 plus.

The emphasis is on the pupils' learning the basic concepts of chemistry through experimentation and observation. Stress is laid on enabling the pupils to develop sufficient skill in simple laboratory techniques and help them appreciate the need for careful observation. These have been the guiding principles for the write up of the laboratory manual and the text. The teacher's guide, as its name implies, is expected to fill in any gaps in the academic make up of the teacher to the extent required in presenting the material to the pupils.

The authors are conscious of the shortcomings and limitations of the material prepared by them. They are aware also that any meaningful improvement in the style and the contents of the books will be possible only after a candid and critical comment on the material, by all those who are genuinely interested in the improvement of standard of chemistry in our schools. The members of the Study Groups have, therefore, great pleasure in sending you copies of the books and invite your comments and constructive suggestions, by way of answers to the questionnaire after a survey of the books.

Any useful suggestions beyond the questionnaire are welcome.

Dept. of Chemistry,
OSMANIA UNIVERSITY,
HYDERABAD-7 A. P.

N. V. SUBBA RAO,
Convener,
Chemistry Study Groups.

QUESTIONNIRE

On the school chemistry laboratory manual, text book and
teacher's guide for classes 6 & 7

N. B. —In case the space provided against each question is insufficient, please attach additional papers. Specific and concrete suggestions for any modification, alteration and improvement of the material will be appreciated.

1. Are the experiments in the laboratory manual simple enough to be easily done by the pupils ?

If not, which experiments, do you feel are difficult ?

How would you like them to be modified ?

2. Are the experiments suited to arouse the curiosity and interest of the pupils ?

Do they lead to an appreciation of the basic principles of chemistry ?

If not, what improvements do you suggest ?

3. In general, are the experiments in the laboratory manuals well planned ?

4. Do you think that the experiments mentioned in the manuals are adequate ?

If not, what additional experiments would you like to be introduced and at what places ?

5. Do you think that the materials required for the experiments would be easily available ?
If not, can you suggest alternatives, using more easily available materials, to bring out the same ideas ?
6. What, in your opinion, is the scope and extent of improvisation that can and may be made to replace the equipment mentioned ?
7. What, in the context of the present condition of our schools, is your opinion about the feasibility of the idea of pupil experimentation ?
8. The material in the text has been prepared, to a large extent, to help the pupils learn chemistry by drawing conclusions from his own observations in the laboratory.
Do you think that this objective has been broadly realised ?
9. Is the language on the whole simple enough for an easy understanding of the subject ?
If not, point out where the language could be simplified.
10. Do you think that the material as developed in the text makes for continued and logical development of ideas and concepts ?
11. Are the contents of each unit within the comprehension of pupil of the age group of 10+ and 12+ ?
If not, is it possible to express the same concept in a simpler way ?

12. Do you think that the sequence of the units is satisfactory? If not, Please suggest how the sequence may be rearranged to enable a better development of the subject.

Give reasons for your rearrangement.

13. Are the diagrams and illustrations suggestive and self-explanatory? State how they may be improved. Do you think that the present illustrations are adequate?

If not, which additional diagrams do you suggest?

14. Do the demonstration experiments mentioned in the teacher's guide bring about the concepts?

If not, what alternative experiments would you suggest?

15. The teacher's guide is intended to enable the teacher to teach the subject confidently and more effectively. It guides him to plan the lesson and provides background information to extend the sphere of his knowledge and answer questions which might be put by intelligent pupils. How far does this teacher's guide serve the purpose for which it is intended?

16. Do you think that the teacher's guide contains sufficient information for the teacher to present the topic effectively?

If not, what additional information may be given and where?

17. The aim of this effort is to introduce the pupils to modern concepts in chemistry and promote a sense of experimentation and inquiry in him. Do you think that this would be achieved by the text and laboratory work presented ?

What alterations would you suggest in this regard ?

18. What other pertinent comments and suggestions would you like to make ?

Signature

Station :

Name :

Date .

Address :

